G.C.E. 'O' Level

CHEMISTRY Matters

APPROVED BY MINISTRY OF EDUCATION
for use from 2007-2011

Marshall Cavendish Education

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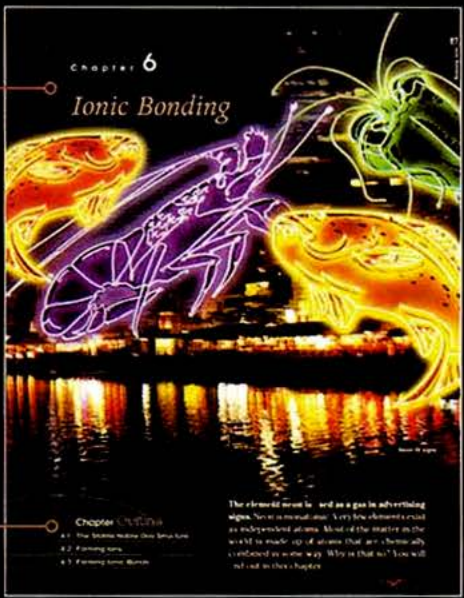
Preface

Chemistry Matters builds on the strength of the highly acclaimed Chemistry Matters for GCE 'O' Level to provide the best possible foundation or success in the study of Chemistry. Designed for effective concept development and reinforcement, this textbook meets the aims and objectives of the Chemistry GCE O-Level Syllabus as specified by the Ministry of Education, Singapore.



Chapter Opener

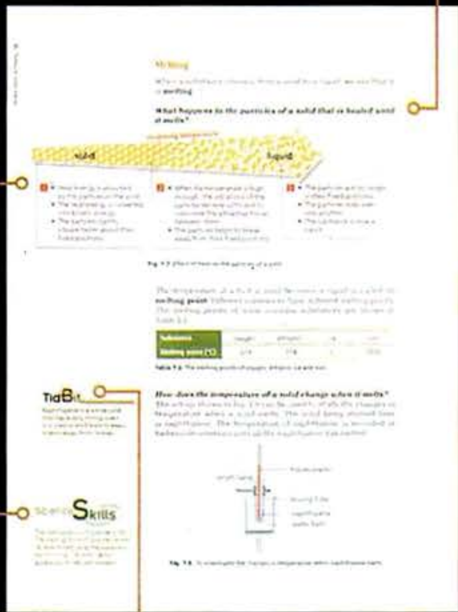
A full-page lead-in to every chapter provides interesting stimulus material and probing questions to set students thinking on how chemistry affects our daily lives.

Chapter Outline

This is an overview of what students will learn in the chapter.

Headers

Headers in the form of questions or short phrases help students to focus on key concepts and syllabus learning outcomes.



Infographics

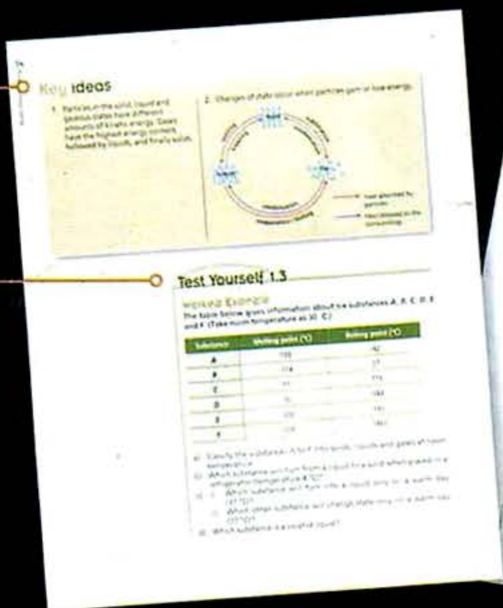
Flowcharts, concept maps and annotated diagrams make chemical processes easier to understand.

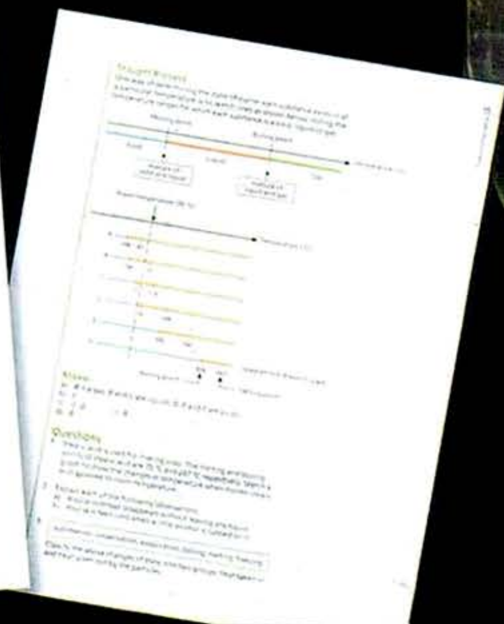
Science Skills

Contains short text or focusing questions designed to help students appreciate how theoretical concepts are derived from experimental investigations. The text may also highlight experimental precautions or guide the student to think about the limitations of the experimental set-up.

Tidbit

Provides extra explanatory or interesting information that bring learning Chemistry out of the classroom.



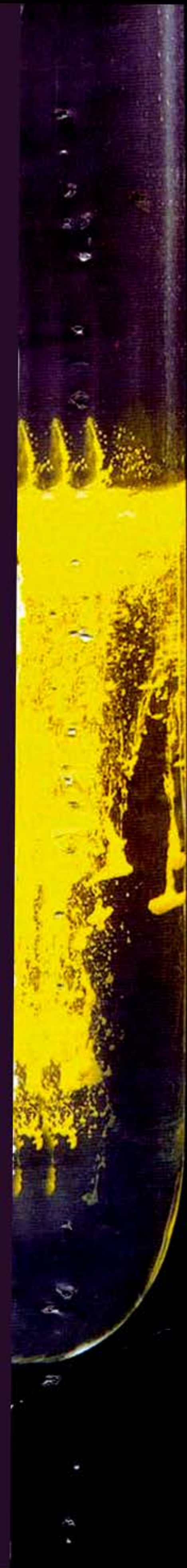


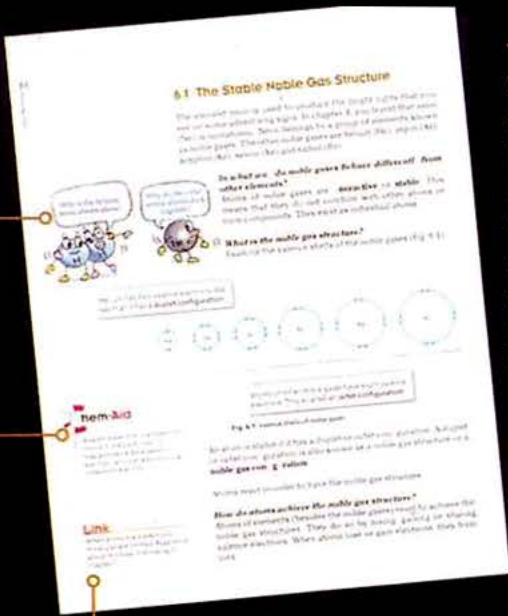
Key ideas

A summary of key concepts at the end of each section provides regular revision and reinforcement.

Test Yourself

Provides worked examples showcasing possible thought processes for problem solving and step by step guidance to build up thinking skills. This is followed by questions that reinforce learning and understanding.





Cartoons

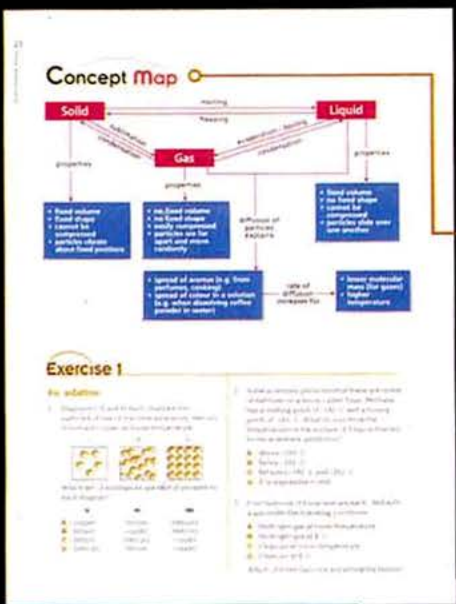
Demonstrate concepts and serve as visual 'hooks' to help students recollect key ideas.

Chem–Aid

These strategically placed pointers help students to understand new concepts by reminding them of prior knowledge and explaining difficult words. They also provide tips and tricks to help students recollect difficult ideas.

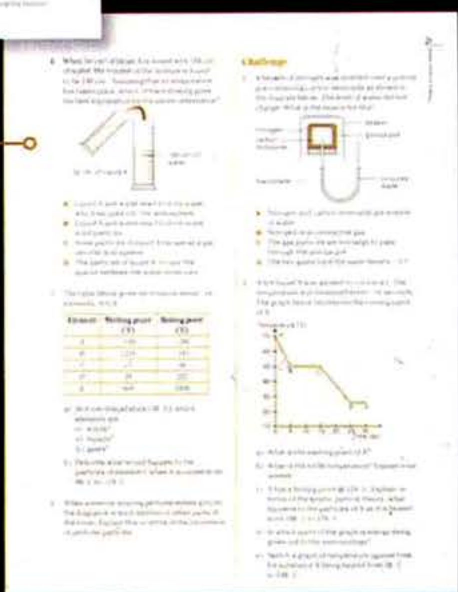
Link

Provides statements and questions that will help students appreciate common chemical concepts that link different chapters.



Concept Map

The end-of-chapter concept map provides a quick summary and links all the key concepts in that chapter.

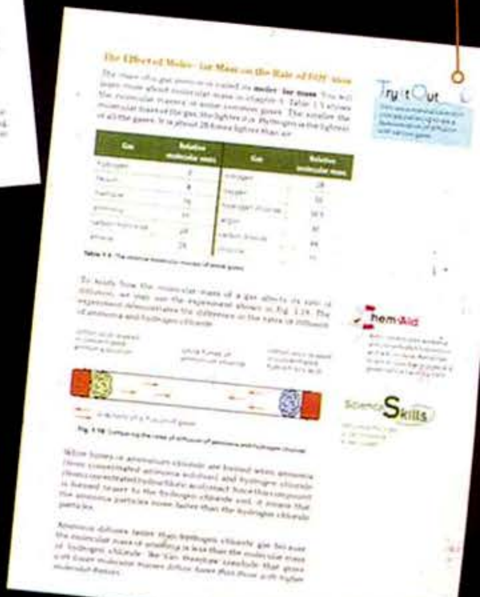


Exercise

This provides a variety of examination-type questions at the end of each chapter to help students integrate and apply what they have learnt. Questions are categorised into 'Foundation' and 'Challenge' sections to suit different levels of testing. Questions that provide practice in interpreting and manipulating different kinds of data are also included.

Try It Out!

Provides hands-on or information technology based activities to help build understanding and stimulate further exploration. A 'mouse' icon indicates that more topic-related enrichment material can be found online. A list of suggested URLs is provided at the back of the book for Internet related activities.



Chemistry Today

Provides excellent opportunities to explore related topics and stimulate class discussion. Contains Critical Thinking questions that help students to develop and hone higher order thinking skills by considering natural phenomena and everyday events in the light of what they have learnt.

Contents

Chapter 1

Kinetic Particle Theory

Everything around you is made up of matter. Living things like your pet cat and even you are made of matter. Non-living things such as a roll of film, popcorn and dust are made of matter too.

Have you ever seen dust 'dancing' in the light? If not, the next time you watch a movie, look carefully at the beam of light from the back of the theatre. The dust moves irregularly in the beam of light that shines from the movie projector. What makes dust move that way? You will learn the answer in this chapter.

Chapter Outline

1.1 | States of Matter

Matter can exist as a solid, a liquid or a gas. These three forms of matter are called the **states of matter**. For example, water (liquid) can exist as ice (solid) or water vapour (gas).

Most substances can exist in each of the three states of matter. Matter can change from one state to another due to changes in temperature and pressure. For example, on freezing, water becomes ice; on boiling, water becomes water vapour.

The three states of matter have very different properties as summarised in Table 1.1.

Property	Solid	Liquid	Gas
Shape	fixed	not fixed	not fixed
Volume	fixed	fixed	not fixed
Compressibility	cannot be compressed	cannot be compressed	can be compressed

Table 1.1 Properties of solids, liquids and gases

Chem-Aid

1. A vapour is a gas that exists at room temperature and pressure (r.t.p.). For example, water vapour is in the air we breathe. Vapour can be given off by a liquid even when the liquid is not boiling.

2. To indicate the state of a substance, we use the state symbols **s** for solid, ***l*** for liquid and **g** for gas or vapour. The state symbol **aq** is used for substances dissolved in water. The state symbol is written within brackets and placed after the *formula* of the substance.

formula of substance
state symbol

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

calcium carbonate	hydrochloric acid	calcium chloride	water	carbon dioxide

name of substance

These state symbols are important when writing chemical equations.

Key ideas

1. Matter is a substance that has mass and occupies space.
2. The three states of matter are solid, liquid and gas.
3. Changes in temperature and pressure can change the state of matter.

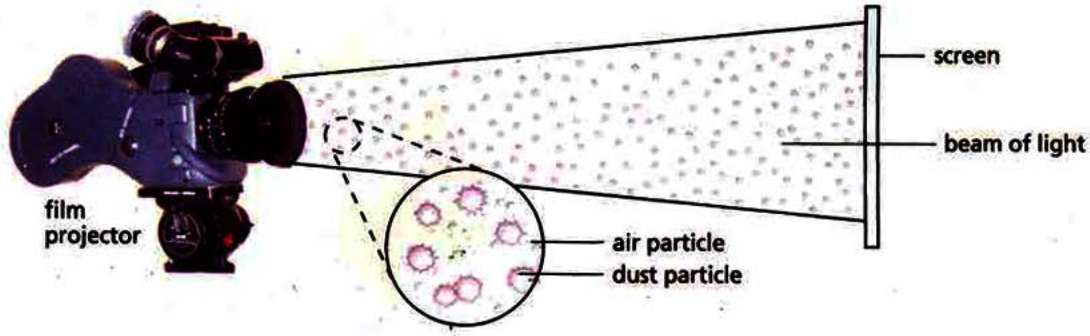
Test Yourself 1.1

Questions

1. Name the different states of matter in which burning candle wax can exist.

2. Mercury is used in laboratory thermometers. At r.t.p. (room temperature and pressure), which state of matter does it exist in?

1.2 | Kinetic Particle Theory



Did you know that air is made up of tiny particles that move around? The 'dancing' dust that you see in a beam of light is actually the result of air particles moving and bumping into dust! Air particles are too small to be seen by our eyes, therefore we can only see the dust moving.

This explanation for 'dancing' dust is based on the **kinetic particle theory**. The kinetic particle theory states that *all matter is made up of tiny particles* and that these particles are *in constant, random motion*. Moving particles have kinetic energy, hence the name kinetic particle theory. We shall now learn more about the kinetic particle theory.

Chem-Aid

The word 'kinetic' refers to motion. For example, 'kinetic energy' is the energy of a moving object.

The kinetic particle theory
- describes the states of matter,
- explains the differences in the properties of solids, liquids and gases,
- explains the changes of state.

'Dancing' dust is evidence that particles of air are moving about constantly. What other everyday observations prove that particles of matter are in constant motion?

The Solid State

According to the kinetic particle theory, the particles of a solid are closely packed in an orderly pattern (Fig. 1.1).

Why does a solid have a fixed shape and a fixed volume?
The particles of a solid are held together by very strong forces of attraction. They cannot move about freely. They only have enough kinetic energy to vibrate and rotate about their fixed positions. For this reason, a solid has a fixed shape.

A solid cannot be compressed since its particles are already very close to one another. Thus, a solid also has a fixed volume.

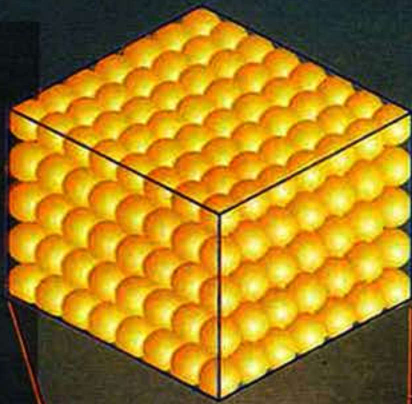


Fig. 1.1 Particles are closely packed together in a solid.

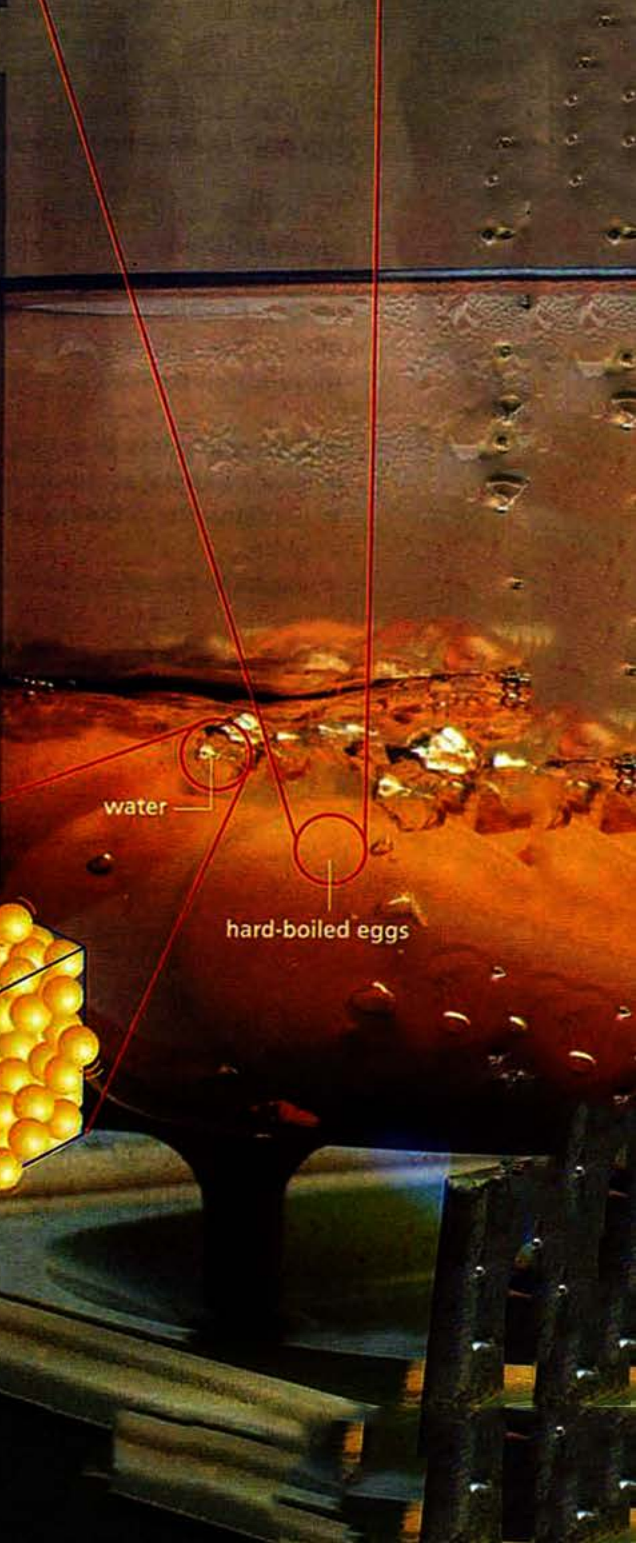
The Liquid State

Look at Fig. 1.2. Compare the arrangement of the particles of a liquid with that of the solid in Fig. 1.1. In a liquid, there is more space between the particles.

Why does a liquid not have a fixed shape?
According to the kinetic particle theory, the forces of attraction between the particles of a liquid are weaker than those in a solid. The particles of a liquid are not held in fixed positions. They are arranged in a disorderly manner and can move freely by sliding over one another. This is why a liquid does not have a fixed shape.

The particles of a liquid have more kinetic energy than particles of the same substance in the solid state.

Why does a liquid have a fixed volume?
The particles of a liquid are farther away from one another than the particles in a solid. However, the particles of a liquid are still packed quite closely together. Thus, a liquid cannot be compressed and has a fixed volume.



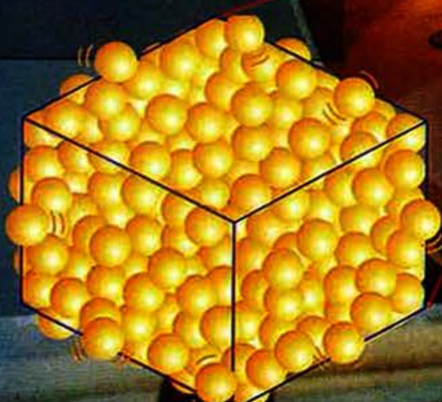


Fig. 1.2 Particles in a liquid are not held in fixed positions.

The Gaseous State

Use Figs. 1.3 and 1.4 to deduce the properties of a gas.

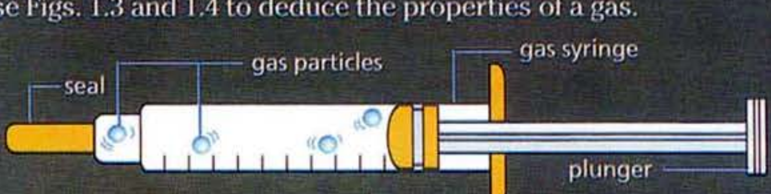


Fig. 1.3 Gas particles are far apart.

Fig. 1.4 Gas particles become closer together.

Why does a gas not have a fixed shape and a fixed volume?

The particles of a gas are spread far apart from one another as shown in Fig. 1.3. This is because the forces of attraction between the particles are very weak.

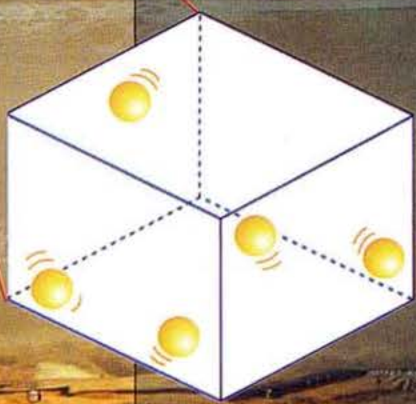
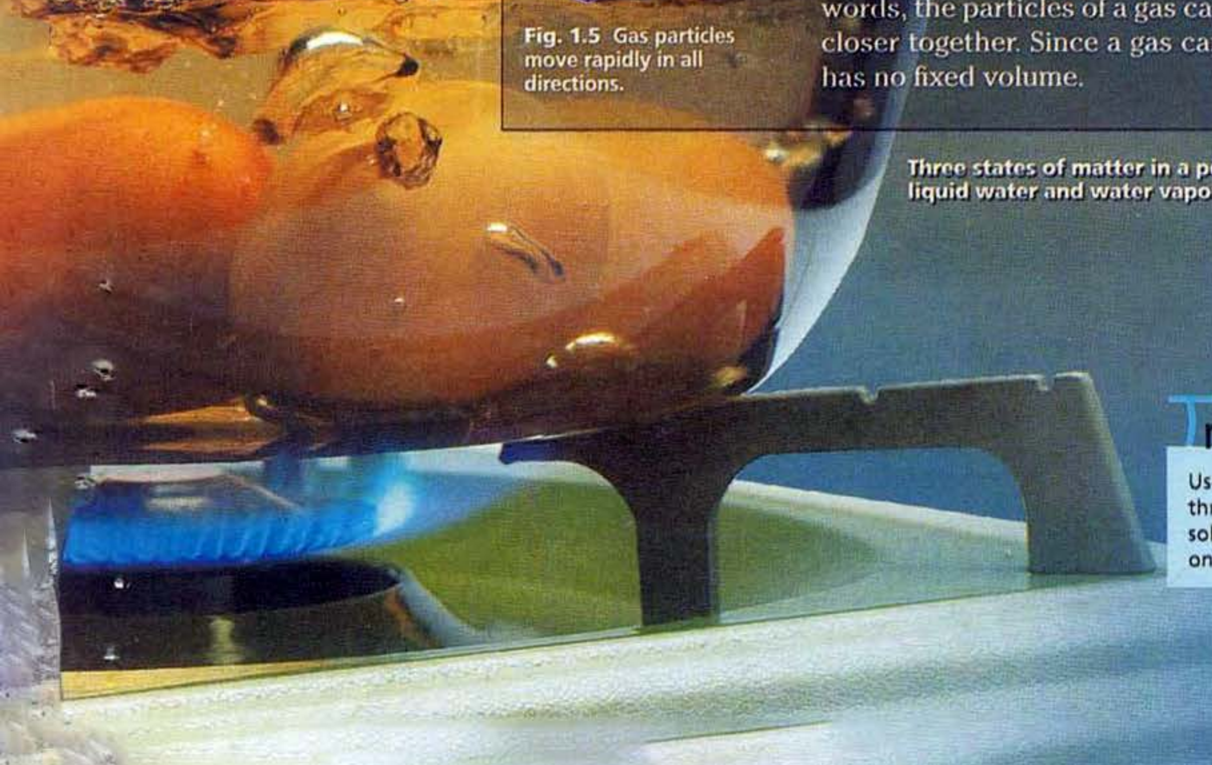


Fig. 1.5 Gas particles move rapidly in all directions.

According to the kinetic particle theory, the particles of a gas have a lot of kinetic energy and are not held in fixed positions. They can move about rapidly in any direction. Thus, a gas has no fixed shape.

Particles of a gas have a lot more space between them as compared to particles of liquids or solids. The large space between the particles of a gas allows the gas to be easily compressed when pressure is applied as shown in Fig. 1.4. In other words, the particles of a gas can be forced to move closer together. Since a gas can be compressed, it has no fixed volume.



Three states of matter in a pot — solid hard-boiled eggs, liquid water and water vapour.

Try it Out

Use the Internet to search for three-dimensional models of solids, liquids and gases based on the kinetic particle theory.

Key ideas

1. The kinetic particle theory states that all matter is made up of tiny particles that are in constant, random motion.
2. The differences in the states of matter can be explained in terms of the kinetic particle theory.

Properties	Solid	Liquid	Gas
Arrangement of particles	• orderly • closely packed 	• disorderly • less closely packed than in a solid 	• disorderly • very far apart 
Attractive forces between particles	very strong	strong	very weak
Kinetic energy of particles	very low	low	high
Particle motion	vibrate and rotate about a fixed position	slide over each other	move about at great speeds

Test Yourself 1.2

Worked Example

Below is a list of four substances at 20 °C.

mercury, oil, water vapour, common salt

Which of the substances
a) does not have fixed shape and volume, and can be compressed?
b) contains the most orderly arrangement of particles?

Thought Process

a) A gas has no fixed shape and no fixed volume. It can be compressed. Identify the substance that is a gas.
b) A solid has the most orderly arrangement of particles compared to a liquid and a gas. Identify the substance that is a solid.

Answer

a) Water vapour
b) Common salt

Questions

1. a) In which state of matter can the particles move most freely?
   b) In which state of matter are the particles closest together?
   c) Sketch a simple diagram to compare the arrangement of the particles in (a) and (b).
2. Using the kinetic particle theory, explain why
   a) liquids have fixed volumes but not fixed shapes.
   b) ice has a higher density than water vapour.

Robert Hooke (1635 -1703)

Robert Hooke was one of the greatest scientists of his time. He proposed the kinetic particle theory that helps to explain the properties of solids, liquids and gases. Robert Hooke developed the kinetic particle theory in order to explain something that he had observed in the laboratory. This is how science develops — someone observes something he or she cannot explain and then carries out experiments to find out more. What was Robert Hooke trying to explain? You may use the Internet to find out.

1.3 | Changes of State and the Kinetic Particle Theory

The physical state of a substance depends on the temperature and pressure of the surroundings. For example, ice changes to water when it is heated to its melting point. Water changes to water vapour when it is heated to its boiling point.

Melting and boiling are examples of changes of state. Other examples of changes of state are freezing and condensation. *Changes of state are reversible*. For example, on boiling, water changes to water vapour; on cooling, water vapour changes back to liquid water.

What causes a substance to change its state?
According to the kinetic particle theory, particles of matter are in constant motion. Thus, they have **kinetic energy**. Gases have the highest energy content, followed by liquids. Solids have the least energy content among the three states.

Chem-Aid
Energy can be changed from one form to another. Heat energy can become the kinetic energy of the particles of matter.

When matter is heated or cooled, the heat taken in or given out causes the kinetic energy of the particles to change. As a result, a substance changes its state.

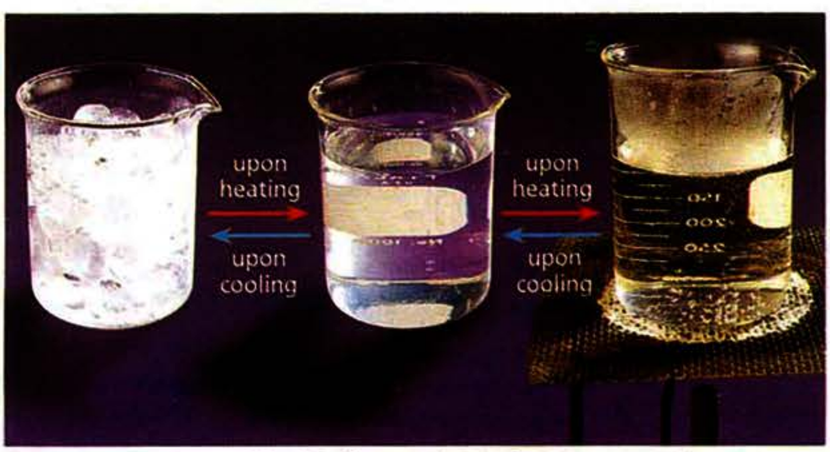
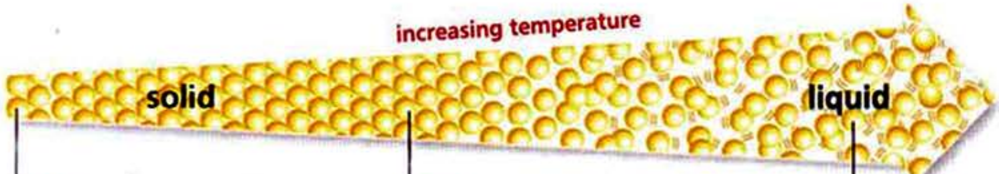


Fig. 1.6 Changes of state of water

Melting

When a substance *changes from a solid to a liquid*, we say that it is **melting**.

What happens to the particles of a solid that is heated until it melts?



1
- Heat energy is absorbed by the particles in the solid.
- The heat energy is converted into kinetic energy.
- The particles start to vibrate faster about their fixed positions.

2
- When the temperature is high enough, the vibrations of the particles become sufficient to overcome the attractive forces between them.
- The particles begin to break away from their fixed positions.

3
- The particles are no longer in their fixed positions.
- The particles slide over one another.
- The substance is now a liquid.

Fig. 1.7 Effect of heat on the particles of a solid

The temperature at which a solid becomes a liquid is called its **melting point**. Different substances have different melting points. The melting points of some common substances are shown in Table 1.2.

Substance	oxygen	ethanol	ice	iron
Melting point (°C)	–219	–114	0	1535

Table 1.2 The melting points of oxygen, ethanol, ice and iron

TidBit

Naphthalene is a white solid that has a very strong smell. It is used in mothballs to keep insects away from clothes.

How does the temperature of a solid change when it melts?
The set-up shown in Fig. 1.8 can be used to study the changes in temperature when a solid melts. The solid being studied here is naphthalene. The temperature of naphthalene is recorded at half-minute intervals until all the naphthalene has melted.

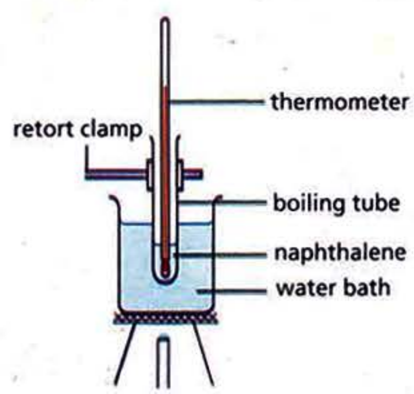


Fig. 1.8 To investigate the changes in temperature when naphthalene melts

ScienceSkills

The melting point of glucose is 142 °C. The melting point of glucose cannot be determined using the apparatus shown in Fig. 1.8. Why? What apparatus can be used instead?

If we plot the graph of temperature against time, we will get the curve shown in Fig. 1.9.

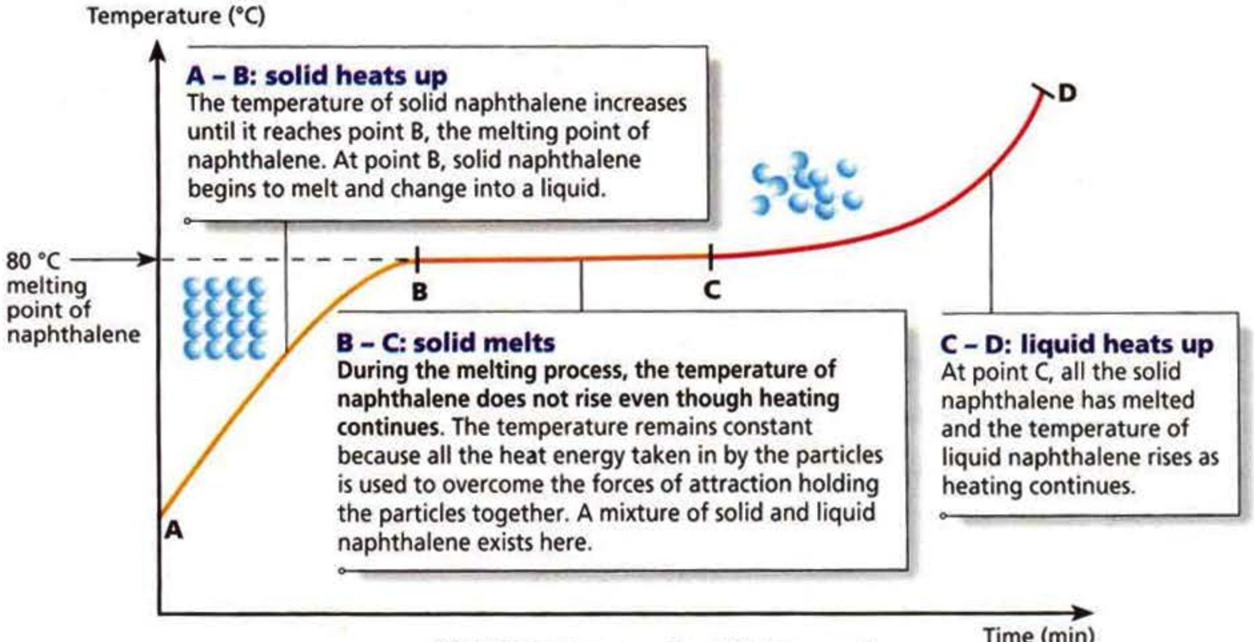


Fig. 1.9 Heating curve of naphthalene

This graph is known as the **heating curve** of naphthalene. A heating curve *shows how the temperature of a solid changes as it is heated to its melting point.*

Freezing

When a substance *changes from a liquid to a solid*, we say that it is **freezing**.

What happens to the particles of a liquid that is cooled until it freezes?

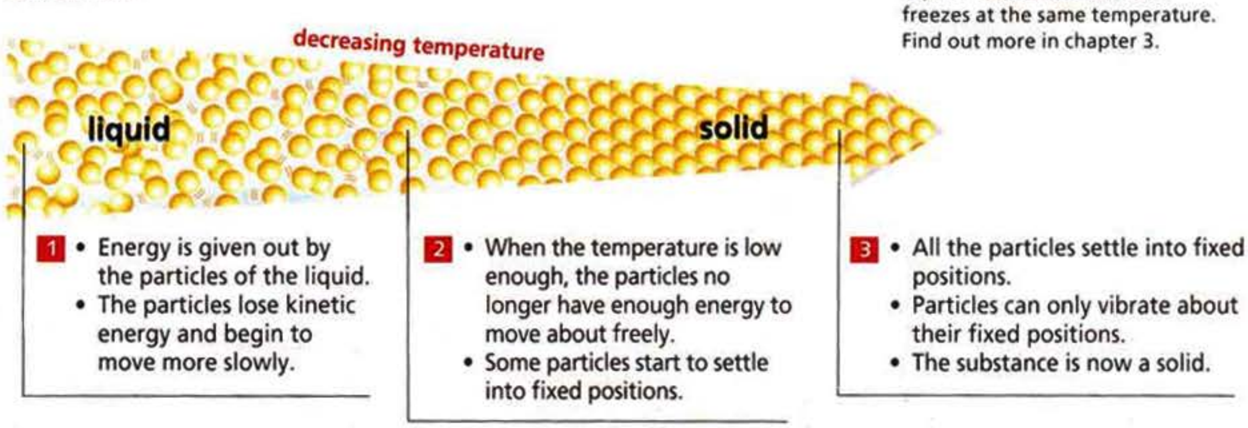


Fig. 1.10 Effect of cooling on the particles of a liquid

Link

A pure substance melts and freezes at the same temperature. Find out more in chapter 3.

The temperature at which a liquid becomes a solid is called its **freezing point**. Heat energy is given out during the cooling process.

How does the temperature of a liquid change when it cools?
We can use naphthalene to study how temperature changes when a liquid cools. Solid naphthalene is first melted. Liquid naphthalene is then allowed to cool in air using the set-up shown in Fig. 1.11.

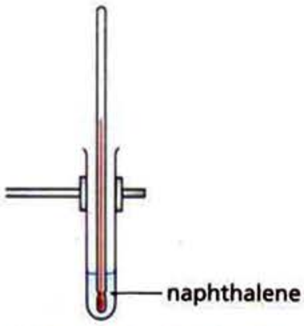


Fig. 1.11 The cooling of naphthalene

The temperature of cooling naphthalene is recorded at half-minute intervals until all the naphthalene has solidified. When temperature is plotted against time, the graph in Fig. 1.12 is obtained. This graph is known as the **cooling curve** of naphthalene. A cooling curve *shows how the temperature of a pure liquid changes as it is cooled to its freezing point.*

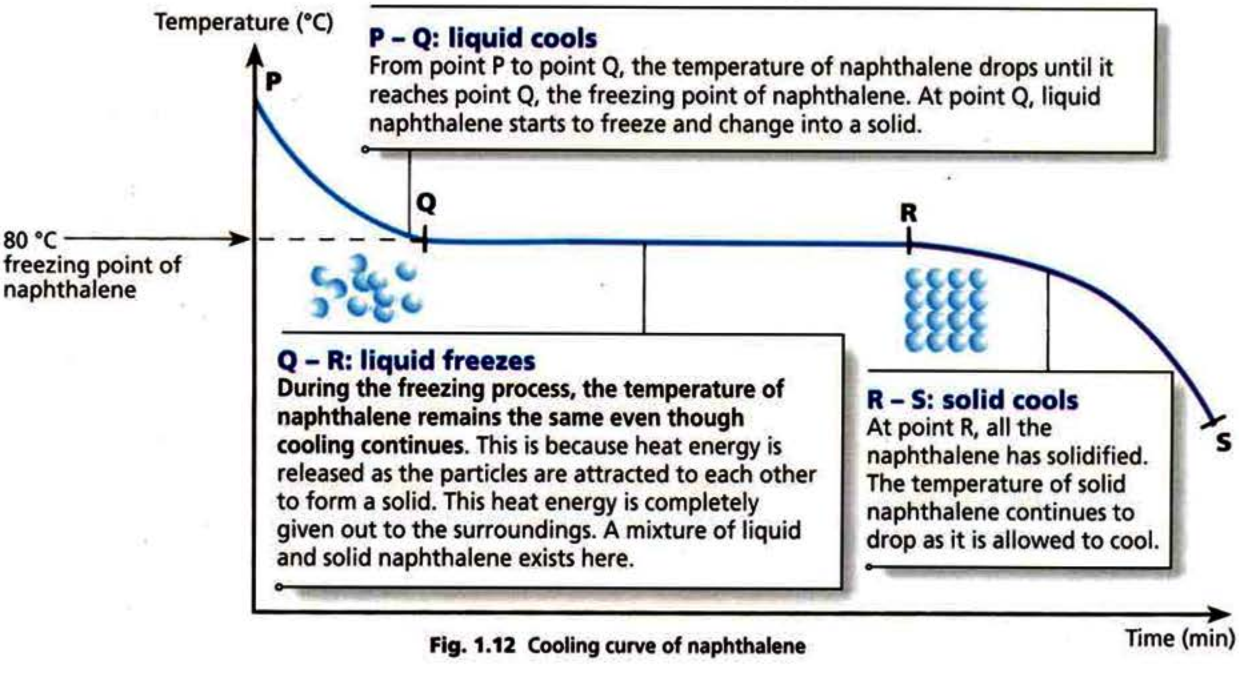


Fig. 1.12 Cooling curve of naphthalene

Boiling

When a substance *changes from a liquid to a gas at the boiling temperature*, we say that it is **boiling**.

What happens to the particles of a liquid that is heated until it boils?

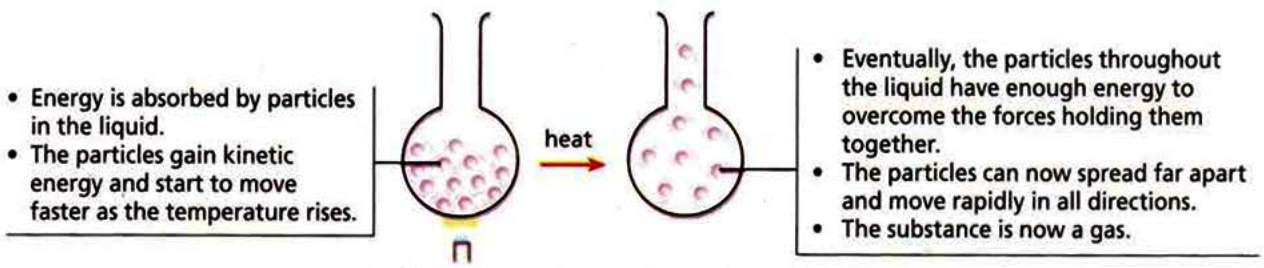


Fig. 1.13 Effect of boiling on the particles of a liquid

The temperature at which a liquid boils to become a gas is called its **boiling point**. Different substances have different boiling points. Table 1.3 shows the boiling points of some common substances at normal atmospheric pressure.

Substance	oxygen	ethanol	water	iron
Melting point (°C)	−183	78	100	3000

Table 1.3 The boiling points of oxygen, ethanol, water and iron

How does the temperature of a liquid change when it boils?
We can use the set-up shown in Fig. 1.14 to determine the boiling point of a liquid at normal atmospheric pressure. In this case, we are determining the boiling point of tetrachloromethane.

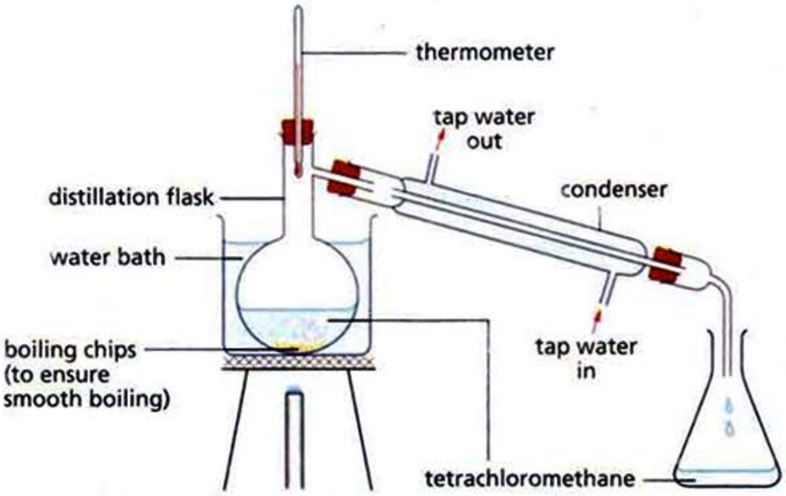


Fig. 1.14 Experiment to determine the boiling point of tetrachloromethane

TidBit

Tetrachloromethane or carbon tetrachloride was used in dry cleaning and fire extinguishers until scientists discovered it is poisonous!

Chem-Aid

This symbol means that a substance is toxic (poisonous). Tetrachloromethane is toxic, and gloves should be worn when handling it. This experiment MUST be carried out in a fume cupboard.

If we record the temperature of tetrachloromethane as it is heated to its boiling point, we can plot a graph like the one shown in Fig. 1.15. Similar graphs are obtained when other liquids are boiled.

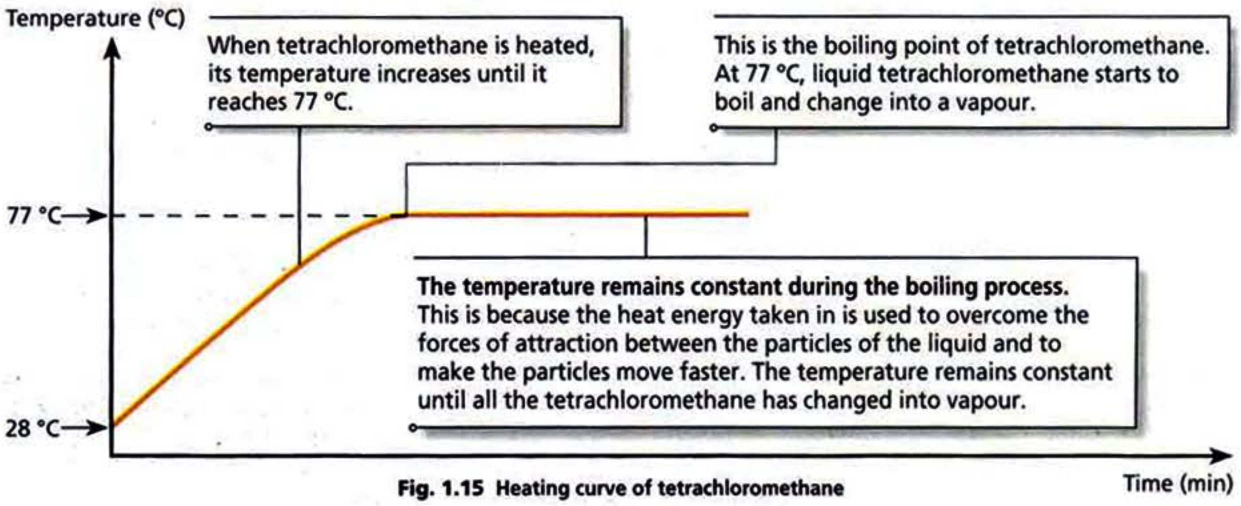
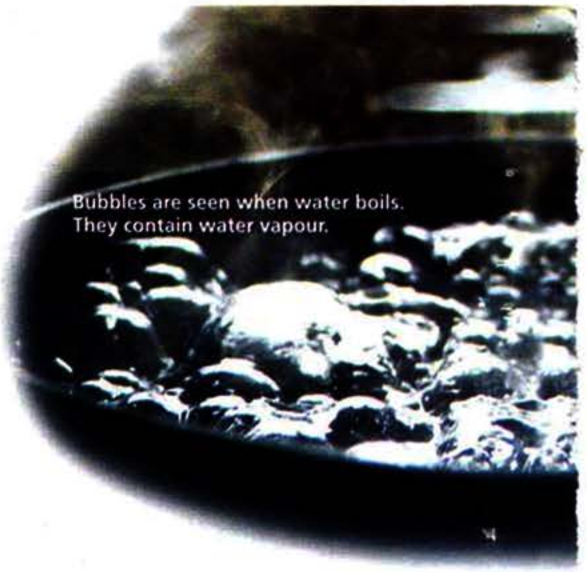


Fig. 1.15 Heating curve of tetrachloromethane



What do you observe when a liquid boils?

When a liquid boils, bubbles of gas are seen. These bubbles are formed when the liquid changes to vapour. The bubbles also consist of other gases dissolved in the liquid. The bubbles rise to the surface and escape into the air.

Evaporation

Why do liquids evaporate?

Sometimes a liquid can turn into a gas at temperatures lower than the boiling point. This process is called **evaporation**. Evaporation occurs because some *particles have enough energy to escape as a gas from the surface of the liquid.*

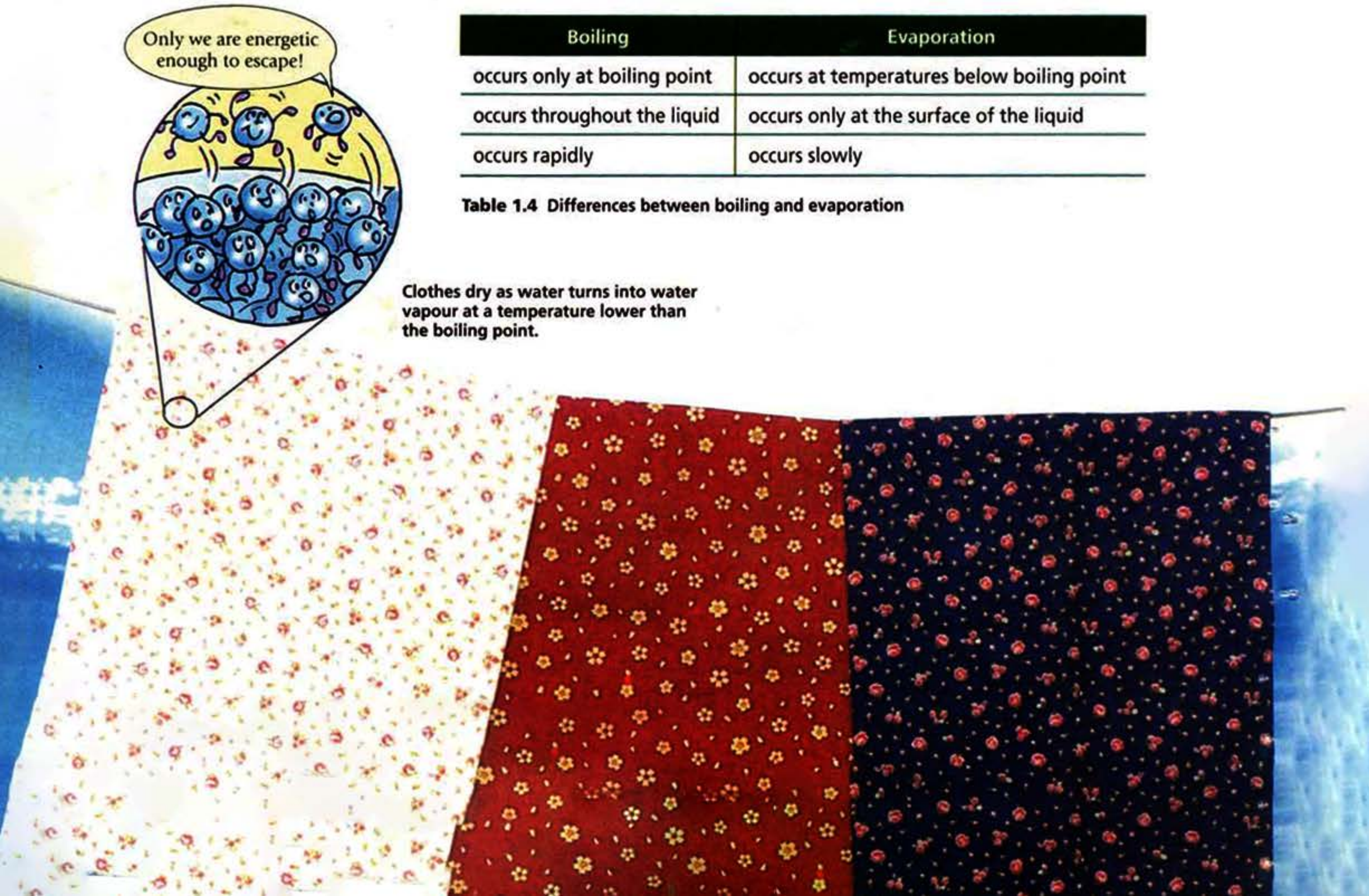
Liquids that evaporate quickly at room temperature are called **volatile liquids**. They usually have boiling points just above room temperature. Petrol and perfumes are examples of volatile liquids.

In what way is evaporation different from boiling?

Both boiling and evaporation involve a liquid changing into a gas, but boiling is different from evaporation in three ways (Table 1.4).

Boiling	Evaporation
occurs only at boiling point	occurs at temperatures below boiling point
occurs throughout the liquid	occurs only at the surface of the liquid
occurs rapidly	occurs slowly

Table 1.4 Differences between boiling and evaporation



Clothes dry as water turns into water vapour at a temperature lower than the boiling point.

Condensation

When a gas is cooled sufficiently, it changes into a liquid. This process is called **condensation**. When water vapour touches a cold surface, condensation occurs and liquid water is obtained (Fig. 1.16).

What happens to the particles of a gas when it condenses?
Think of condensation as the reverse of boiling. Heat energy is given out during condensation. As the temperature drops, the gas particles lose energy and move more slowly. Eventually, the movement of the particles becomes slow enough for the gas to change into a liquid.

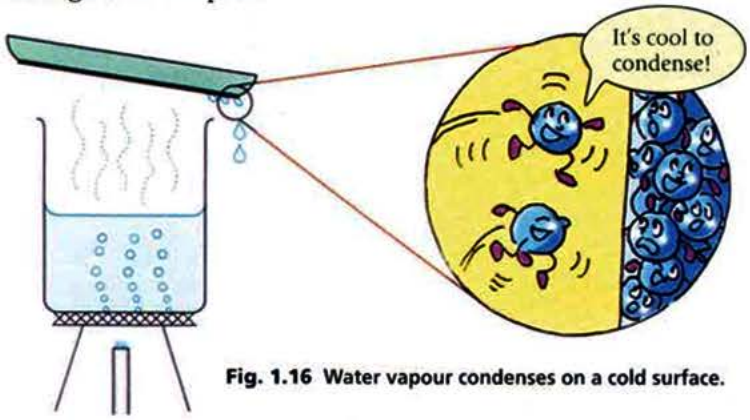


Fig. 1.16 Water vapour condenses on a cold surface.

Try it Out

How could you use condensation to help you survive in a desert?
You cannot live without water. If you were lost in a desert, what would you do? Design your own apparatus for obtaining water. Use some or all of the following materials.

- 1 container
- 1 plastic sheet
- 1 collecting vessel
- 1 small plant
- some sand
- some stones

Once your apparatus is ready, try it out! Observe what happens, measure the amount of water collected and record the values in a notebook. What changes can you make to the set-up to obtain more water?

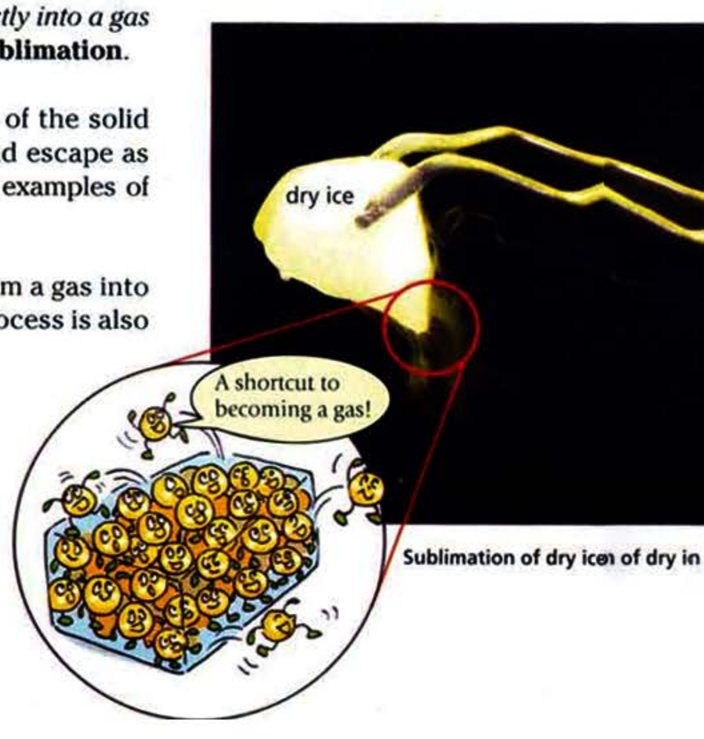
Sublimation

Can a solid change into a gas without turning into a liquid first?
When dry ice (solid carbon dioxide) is exposed to temperatures higher than –78 °C, it turns directly into carbon dioxide gas without melting. Like dry ice, some solids *change directly into a gas without going through the liquid state*. This is called **sublimation**.

Sublimation occurs because particles at the surface of the solid have enough energy to break away from the solid and escape as a gas. Iodine and ammonium chloride are two other examples of solids that sublime.

Substances that sublime may also change directly from a gas into a solid without going through the liquid state. This process is also called condensation.

In what way is sublimation useful?
Dry ice is used for industrial refrigeration and transporting frozen foods. It is especially good for refrigerating food such as ice cream and meat because it keeps them very cold and it changes directly into a gas without leaving any liquid behind.



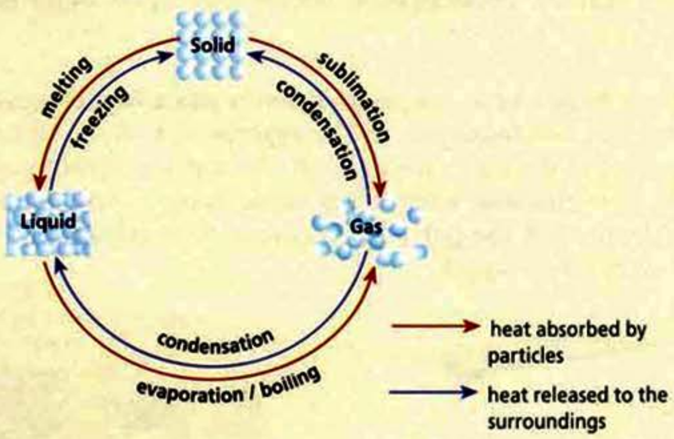
Sublimation of dry icen of dry in

Key ideas

1. Particles in solids, liquids and gases have different amounts of kinetic energy.

   Gases have the highest energy content, followed by liquids, and finally solids.

2. Changes of state occur when particles gain or lose energy.



Test Yourself 1.1

Worked Example

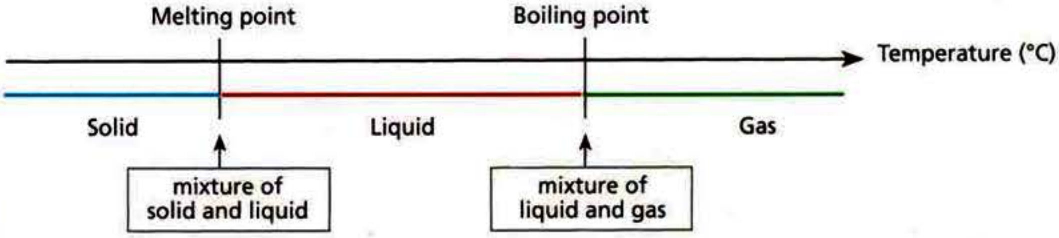
The table below gives information about six substances, *A*, *B*, *C*, *D*, *E* and *F*. (Take room temperature as 30 °C.)

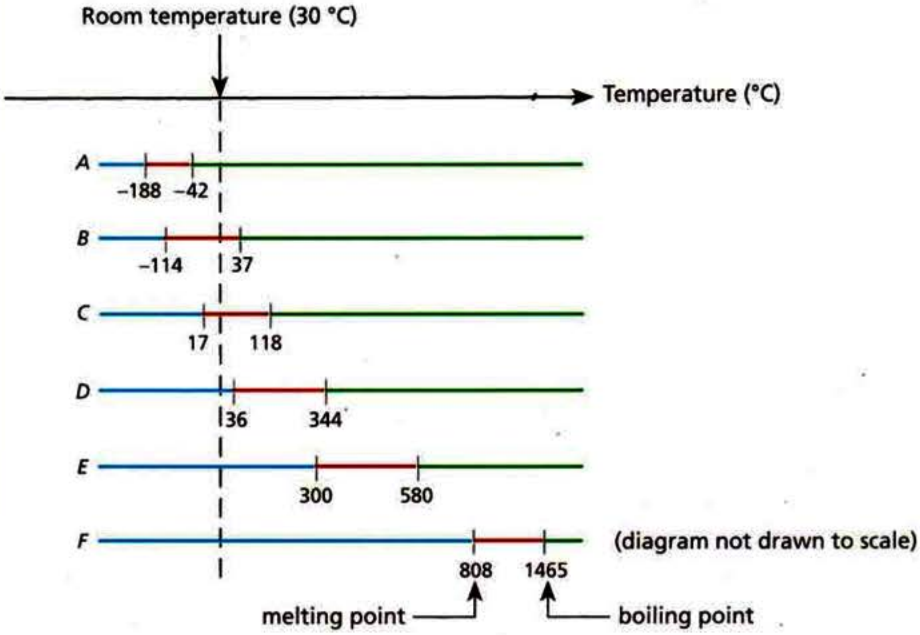
Substance	Melting point (°C)	Boiling point (°C)
A	−188	−42
B	−114	37
C	17	118
D	36	344
E	300	580
F	808	1465

a) Classify the substances *A* to *F* into solids, liquids and gases at room temperature.
b) Which substance will turn from a liquid to a solid when placed in a refrigerator (temperature 4 °C)?
c) i) Which substance will turn into a liquid only on a warm day (37 °C)?
   ii) Which other substance will change state only on a warm day (37 °C)?
d) Which substance is a volatile liquid?

Thought Process

One way of determining the state of matter each substance exists in at a particular temperature is to sketch lines as shown below, noting the temperature ranges for which each substance is a solid, liquid or gas.





Answer

a) *A* is a gas. *B* and *C* are liquids. *D*, *E* and *F* are solids.
b) *C*
c) i) *D* ii) *B*
d) *B*

Questions

1. Stearic acid is used for making soap. The melting and boiling points of stearic acid are 70 °C and 287 °C respectively. Sketch a graph to show the changes in temperature when molten stearic acid is cooled to room temperature.

2. Explain each of the following observations.
   a) A solid mothball disappears without leaving any liquid.
   b) Your skin feels cold when a little alcohol is rubbed on it.

3. | sublimation, condensation, evaporation, boiling, melting, freezing |
   |---|

   Classify the above changes of state into two groups: heat taken in and heat given out by the particles.

1.4 | Diffusion

Diffusion in Gases

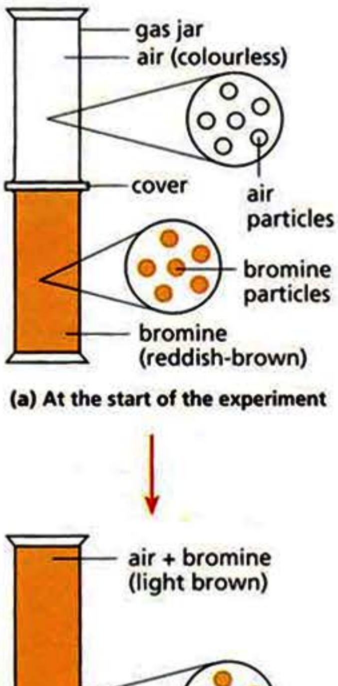
When a bottle of perfume is left open for some time, the scent of the perfume soon spreads throughout the entire room. Similarly, if your mother is cooking curry in the kitchen, you will soon be able to detect the smell of spices in every room.

How does smell spread?
Millions of tiny gas particles escape from the surface of perfumes and spices. These particles can move around freely until they reach your nose, where they are detected as aromas. The spreading of aromas is another observation that proves that particles of gases are constantly moving. *The process by which particles move freely to fill up any available space* is called **diffusion**.

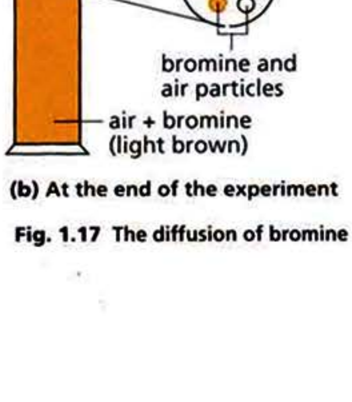
Handle bromine with care. It is poisonous.

We can demonstrate diffusion in the laboratory by placing a gas jar of air over a gas jar containing bromine vapour (Fig. 1.17). A cover is initially placed between the two gas jars to separate the gases. Air is colourless. Bromine vapour is reddish-brown and is heavier than air. A few minutes after the cover is removed, the gas in both gas jars looks the same.

This happens because both air and bromine are made up of tiny particles moving at random. The bromine particles diffuse (spread out) into the spaces between the air particles while the air particles diffuse into the spaces between the particles of bromine. When the gas looks the same in both gas jars, it means that the *particles* of each gas *are evenly spread throughout* both gas jars. We say that a **homogenous mixture** of air and bromine is formed.



(a) At the start of the experiment



(b) At the end of the experiment

Fig. 1.17 The diffusion of bromine

Chem-Aid

A homogenous mixture looks the same throughout.

Do all gases diffuse at the same rate?
A balloon filled with hydrogen gas will shrink faster than a balloon filled with air. The rubber that is used to make balloons contains millions of tiny holes. The smaller hydrogen particles can diffuse more quickly through the holes than the larger air particles, which consist mainly of oxygen, nitrogen and carbon dioxide. Thus, *gas particles of different sizes diffuse at different rates.*

Besides being smaller, the hydrogen particles are also lighter than the other air particles. *Gases of different masses diffuse at different rates.*

The Effect of Molecular Mass on the Rate of Diffusion

The *mass of a gas particle* is called its **molecular mass**. You will learn more about molecular mass in chapter 9. Table 1.5 shows the molecular masses of some common gases. The smaller the molecular mass of the gas, the lighter it is. Hydrogen is the lightest of all the gases. It is about 20 times lighter than air.

Gas	Relative molecular mass	Gas	Relative molecular mass
hydrogen	2	nitrogen	28
helium	4	oxygen	32
methane	16	hydrogen chloride	36.5
ammonia	17	argon	40
carbon monoxide	28	carbon dioxide	44
ethene	28	chlorine	71

Table 1.5 The relative molecular masses of some gases

To study how the molecular mass of a gas affects its rate of diffusion, we may use the experiment shown in Fig. 1.18. The experiment demonstrates the difference in the rates of diffusion of ammonia and hydrogen chloride.

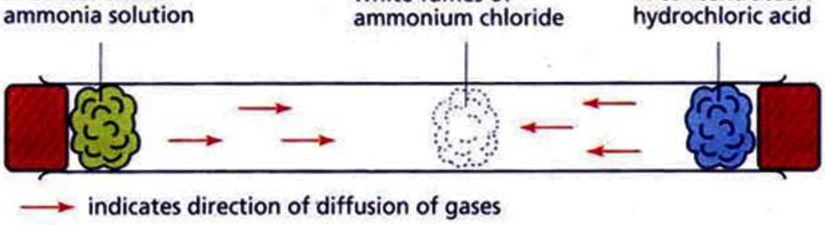


Fig. 1.18 Comparing the rates of diffusion of ammonia and hydrogen chloride

White fumes of ammonium chloride are formed when ammonia (from concentrated ammonia solution) and hydrogen chloride (from concentrated hydrochloric acid) react. Since this compound is formed nearer to the hydrogen chloride end, it means that the ammonia particles move faster than the hydrogen chloride particles.

Ammonia diffuses faster than hydrogen chloride gas because the molecular mass of ammonia is less than the molecular mass of hydrogen chloride. We can therefore conclude that *gases with lower molecular masses diffuse faster than those with higher molecular masses.*

Try it Out

Use the Internet to search for and view a demonstration of diffusion with various gases.

Chem-Aid

Both concentrated ammonia and concentrated hydrochloric acid are corrosive. Remember to put on your goggles and gloves before handling them.

Science Skills

Why must the tube
a) be horizontal?
b) be stoppered?

Science Skills

Suggest possible reasons why distilled water is used instead of tap water in the experiment in Fig. 1.19.

Diffusion in Liquids

Diffusion also takes place in liquids. If a small crystal of potassium manganate(VII) is introduced into a beaker of distilled water, it dissolves to form a deep purple solution, which settles at the bottom of the beaker. Diffusion slowly takes place until the solution becomes uniformly purple (Fig. 1.19).

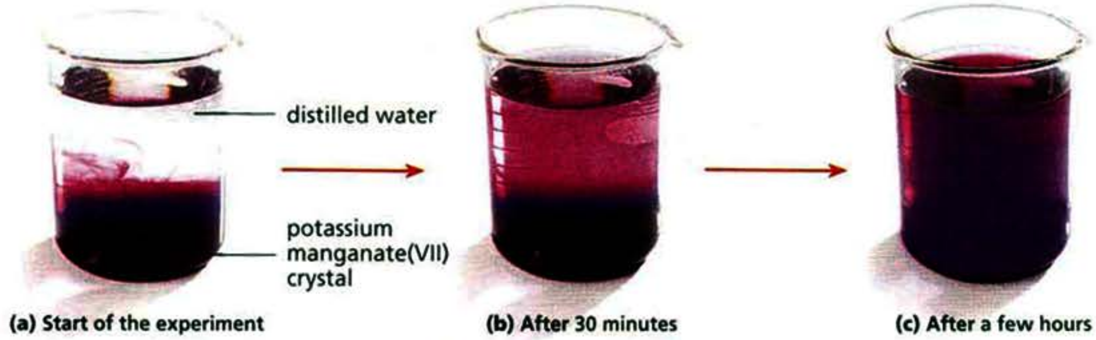


Fig. 1.19 The diffusion of potassium manganate(VII) in water

This illustrates an important point: **diffusion** is *the movement of particles from a region of higher concentration to a region of lower concentration*. This applies to gases as well.

Effect of Temperature on the Rate of Diffusion

What would you observe if the experiment in Fig. 1. 19 was repeated using boiling water? The solution would become uniformly purple within a much shorter time. The rate of diffusion increases as the temperature of the solution increases.

An inflated balloon shrinks faster on a warm day than on a cold day. A lump of sugar dissolves faster in hot water than in cold water. All these examples show that *the higher the temperature, the faster the rate of diffusion*.

How do you account for these observations?
Particles gain more energy as the temperature increases. They can move faster and this increases the rate of diffusion.

Hot water turns brown much faster than cold water when coffee powder is added. Can you use the movement of particles to explain why?

Key ideas

1. Gases with lower molecular masses diffuse faster than gases with higher molecular masses.
2. Diffusion is the movement of particles from a region of higher concentration to a region of lower concentration.
3. The relative rates of diffusion in
   - liquids: slow diffusion
   - gases: rapid diffusion

Test Yourself 1.4

Worked Example

When you stir a lump of sugar in water, the sugar dissolves and disappears. Explain the process of dissolving in terms of the movement of particles (i.e. the kinetic particle theory).

Answer

The particles of sugar move from a region where they are in higher concentration (sugar lump) to where they are in lower concentration (water). The particles of sugar diffuse (slowly move) to fill up any available space in between the water particles. This happens until the particles of sugar and water are evenly mixed.

Try it Out

Plan an experiment to investigate how the rate of diffusion of a drop of ink in water changes with temperature.
Consider carefully the following:
- the measurements involved,
- the apparatus required for this experiment,
- the precautions taken,
- the observations and conclusions to be made

Questions

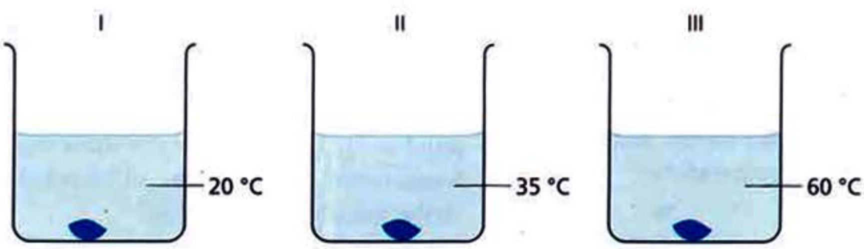
Gas	chlorine	nitrogen	sulphur dioxide	carbon dioxide
Molecular mass	71	28	64	44

a) Which of the gases shown in the table above would diffuse
i) most rapidly?
ii) most slowly?

b) Arrange the gases according to their rates of diffusion (from slowest to fastest).

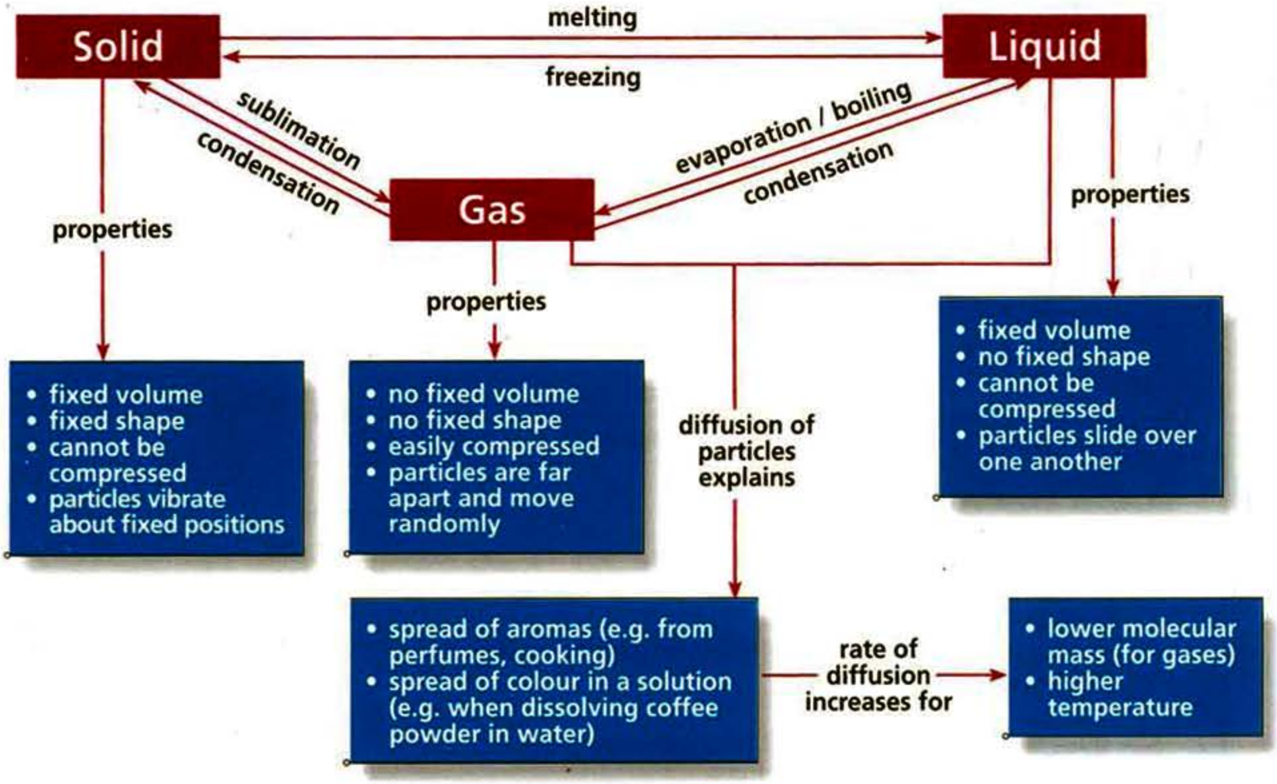
2. Give **three** examples from our daily life to show how important it is for gases to diffuse quickly.

3. Three beakers labelled I, II and III, were each filled with 50 cm^3 of distilled water. The temperature of the water was set at 20 °C, 35 °C and 60 °C respectively. 1 g of copper(II) sulphate crystal was added to each beaker.



In which beaker did copper(II) sulphate take the (i) shortest time, and (ii) longest time to completely diffuse throughout the water? Explain your answer.

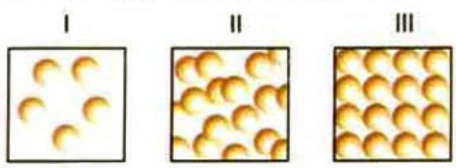
Concept Map



Exercise 1

Foundation

1. Diagrams I, II and III each illustrates the particles of one of the three substances, mercury, helium and copper, at room temperature.



Which set of substances are best illustrated by each diagram?

	I	II	III
A	copper	helium	mercury
B	helium	copper	mercury
C	helium	mercury	copper
D	mercury	helium	copper

2. Some scientists predicted that there are rivers of methane on a moon called Titan. Methane has a melting point of –182 °C and a boiling point of –161 °C. What do you think the temperature on the surface of Titan is that led to the scientists' prediction?

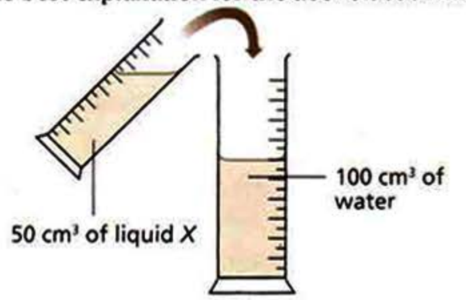
   A Above –161 °C.
   B Below –182 °C.
   C Between –182 °C and –161 °C.
   D It is impossible to tell.

3. Four balloons of equal size are each filled with a gas under one of the following conditions:

   A Hydrogen gas at room temperature.
   B Hydrogen gas at 0 °C.
   C Clean air at room temperature.
   D Clean air at 0 °C.

   Which of these balloons will shrink the fastest?

4. When 50 cm^3 of liquid *X* is mixed with 100 cm^3 of water, the volume of the mixture is found to be 146 cm^3. Assuming that no evaporation has taken place, which of the following gives the best explanation for the above observation?



A Liquid *X* and water react to form a gas, which escapes into the atmosphere.'
B Liquid *X* and water react to form some solid particles.
C Some particles of liquid *X* escape as a gas into the atmosphere.
D The particles of liquid *X* occupy the spaces between the water molecules.

5. The table below gives information about five elements, *A* to *E*.

Element	Melting point (°C)	Boiling point (°C)
A	−189	−186
B	−219	−183
C	−7	58
D	29	222
E	660	2450

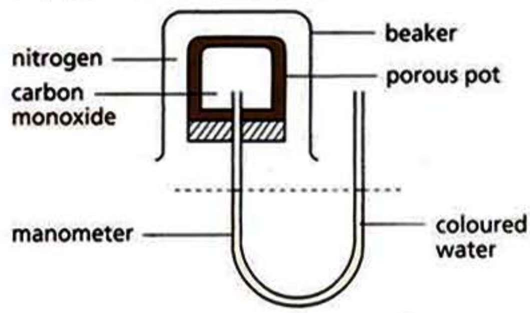
a) At room temperature (30 °C), which element(s) is/are
i) solid(s)?
ii) liquid(s)?
iii) gas(es)?

b) Describe what would happen to the particles of element *C* when it is cooled from 80 °C to −10 °C.

6. When someone who has put on perfume enters a room, the fragrance is soon smelled in other parts of the room. Explain this in terms of the movement of perfume particles.

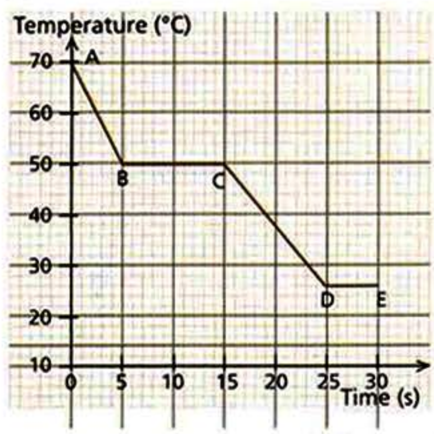
Challenge

1. A beaker of nitrogen was inverted over a porous pot containing carbon monoxide as shown in the diagram below. The level of water did not change. What is the reason for this?



A Nitrogen and carbon monoxide are soluble in water.
B Nitrogen is an unreactive gas.
C The gas particles are too large to pass through the porous pot.
D The two gases have the same density. (*C*)

2. A hot liquid *X* was allowed to cool in air. The temperature was measured every five seconds. The graph below represents the cooling curve of *X*.



a) What is the melting point of *X*?

b) What is the room temperature? Explain your answer.

c) *X* has a boiling point of 128 °C. Explain, in terms of the kinetic particle theory, what happens to the particles of *X* as it is heated from 100 °C to 150 °C.

d) In which parts of the graph is energy being given out to the surroundings?

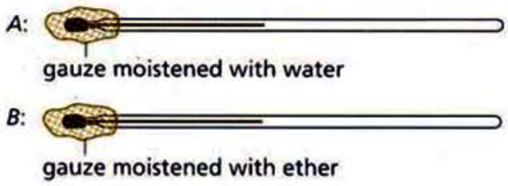
e) Sketch a graph of temperature against time for substance X being heated from 30 °C to 140 °C.

3. Explain the following observations.

   a) When solid ammonium chloride is heated in a test tube, it 'disappears' and then reforms on the colder part of the test tube. Explain why this is so.

   b) If the heating is continued, ammonia is detected first at the mouth of the test tube, followed by hydrogen chloride.

4. Two thermometers, *A* and *B*, are wrapped in pieces of gauze. The gauze around thermometer *A* is moistened with water and the gauze around thermometer *B* is moistened with ether. Ether is a volatile liquid. At the start of the experiment, both thermometers record the same temperature.

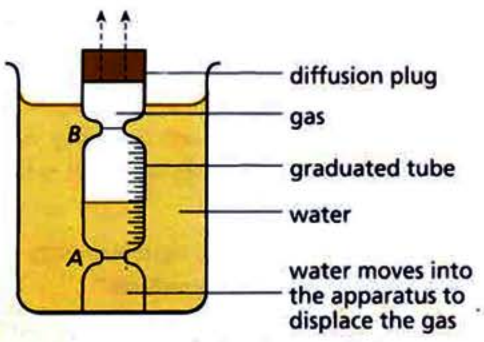


   a) What is meant by 'volatile'?

   b) Suggest why, after 10 minutes, thermometer *B* records a lower temperature than thermometer *A*.

   c) Ether is a flammable liquid. State an important precaution you would take when using ether in this experiment.

5. The diagram below shows an apparatus for measuring the rates of diffusion of gases at a constant temperature. The rate of diffusion is measured in terms of the time taken for the water level to rise from *A* to *B* as the gas diffuses out of the diffusion plug. The table shows the results of this experiment.



Gas	Time (seconds)
carbon monoxide	119
chlorine	190
methane	90
oxygen	127

   a) Explain what is meant by 'diffusion'.

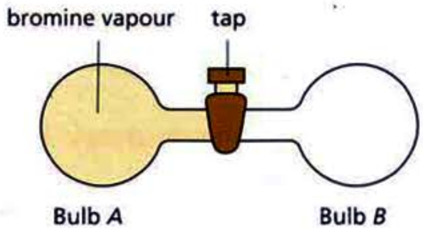
   b) Which gas diffuses slowest? Why?

   c) Name a gas that will diffuse slower than any of the gases shown in the table.

   d) Suggest a reason why hydrogen chloride gas cannot be used in this experiment.

   e) Suggest why ethene is expected to take the same time to diffuse as carbon monoxide.

6. The apparatus shown below was used to investigate the diffusion of bromine.



Experiment 1: Bulb *B* was filled with air. When the tap was opened, the bromine slowly diffused to fill both bulbs.
Experiment 2: The air was pumped out of bulb *B*. When the tap was opened, bromine diffused very rapidly.

Explain these observations in terms of the kinetic particle theory.

Chapter 2

Measurement and Experimental Techniques

Chapter Outline

What do chemists do? In a word — experiment. Through experiments, chemists find out more about the world and create new, useful substances. We see the work of chemists in everyday life. They make the painkillers we take for headaches; the fuel used in racing cars; the plastic called PVC, which is used as artificial leather in furniture; the artificial sweeteners in low-calorie soft drinks and the stretchable material, Lycra™, that is woven into clothes.

Chemists design and perform experiments to test their ideas or to find out more about substances. Their experiments must give trustworthy results. Therefore, it is important that chemists use tools that take accurate measurements.

You too will need to carry out experiments in your school laboratory. Let us find out which physical quantities are usually measured in the experiments you will conduct. You will also learn how these quantities can be measured and which measuring instruments are most suitable for measuring them.

2.1 | Measuring Time

The International System of Units (S.I. units) is a system of different units used to measure quantities of different things. In order to communicate scientific information easily, scientists all over the world use S.I. units as a common standard for their measurements.

In the laboratory, a **stopwatch** or **stopclock** is used to measure time. The S.I. unit for time is the **second (s)**. Other units, such as the **minute (min)** and the **hour (h)**, are used to measure longer intervals of time. Two types of stopwatches are shown in Fig. 2.1.

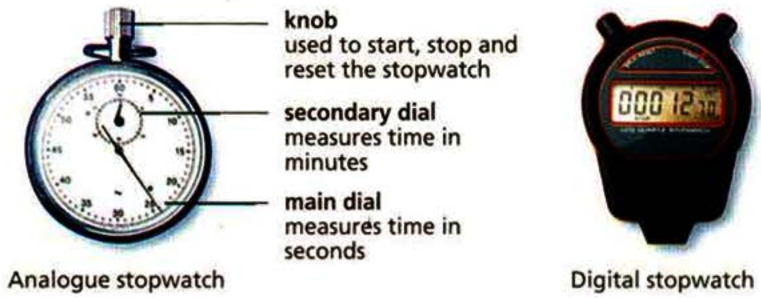


Fig. 2.1 Apparatus for measuring time

2.2 | Measuring Temperature

In the school laboratory, the **mercury thermometer** can be used to measure temperature (Fig. 2.2). The S.I. unit for temperature is the **Kelvin (K)**. However, another unit for temperature, the **degree Celsius (°C)**, is also commonly used.

The mercury thermometer measures temperature from –10 °C to 110 °C. Each division on the scale of this thermometer represents one degree Celsius (1 °C). The accuracy of this type of thermometer is ±0.5 °C.

A temperature sensor can be connected to a data logger to measure temperature. This gives more accurate readings of temperature than a mercury thermometer. Data logging is useful for recording temperature changes continuously over a period of time. You can use a data logger to record and study the changing temperature of a substance that is being heated or cooled. Data loggers are also commonly used for taking measurements outdoors, e.g. humidity of different environments and pH of different sources of water.

Chem-Aid

The Kelvin and Celsius temperature scales are related to each other by the formula:
Temperature in K
= temperature in °C + 273
For example, 0 °C is 273 K.

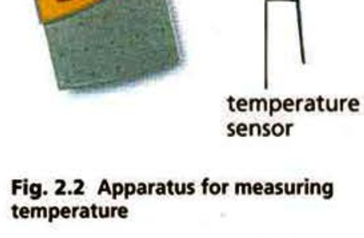


Fig. 2.2 Apparatus for measuring temperature

2.3 | Measuring Mass

The mass of a substance is measured with a **beam balance** or an **electronic balance** (Fig. 2.3). The S.I. unit for mass is the **kilogram (kg)**. Often, very small masses of substances need to be measured in chemistry experiments. In these cases, an electronic balance, with an accuracy of up to 0.01 g (two decimal places), is used for precise and fast weighing.

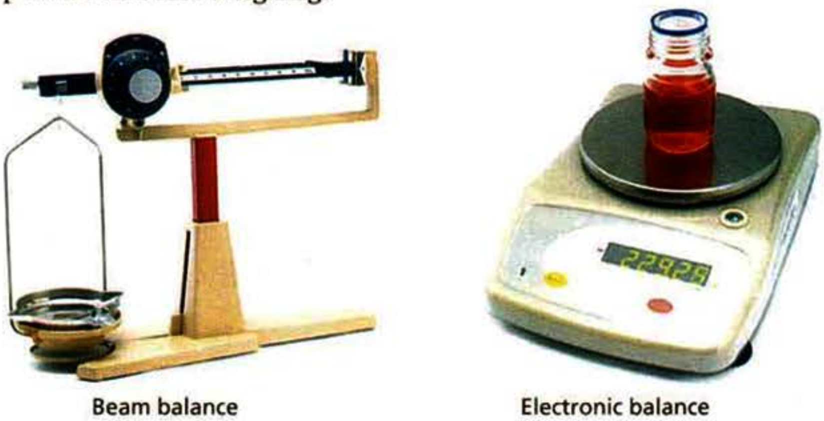


Fig. 2.3 Apparatus for measuring mass

Using a mass balance, a Chinese physician carefully measures out amounts of different herbs to make traditional medicine. When making modern medicines, chemists often use mass balances that are many times more sensitive.

Chem-Aid

The unit for mass, 'kg', is written in lower case. Capital 'K' means Kelvin.

2.4 | Measuring Volume

Scientific experiments often involve liquids. Fig. 2.4 shows some laboratory apparatus commonly used to measure the volume of a liquid. The different apparatus have different degrees of accuracy. Which apparatus you choose for an experiment depends on the volumes you are measuring and how exact you need the volumes to be.

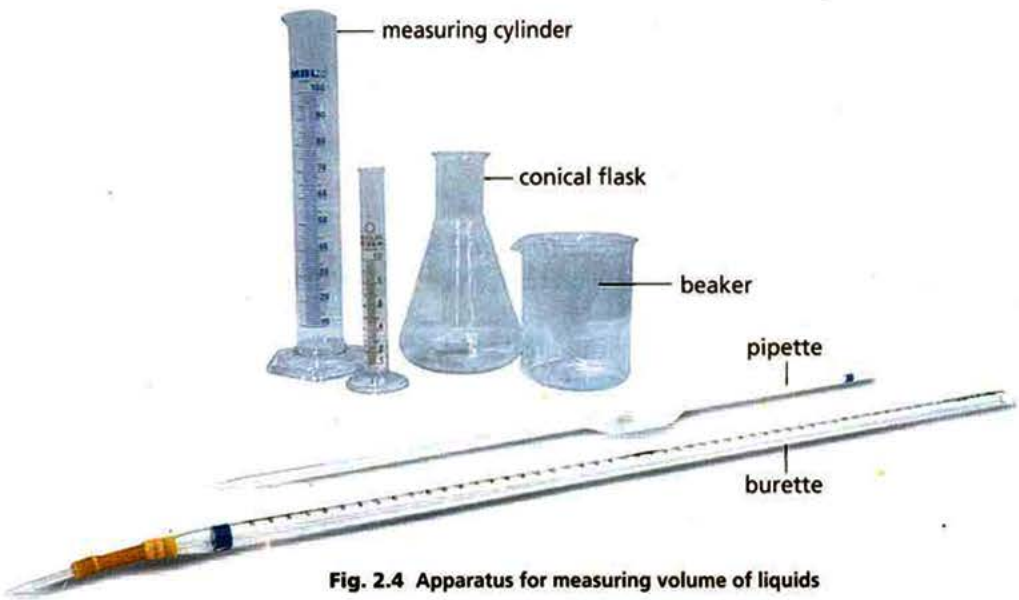


Fig. 2.4 Apparatus for measuring volume of liquids

TidBit

A medical technologist measures out a number of equal volumes of blood using a micropipette. The blood is being tested for the hepatitis virus. Medical technologists handle very tiny volumes of liquid, as small as a millionth of a litre.

It would be impossible to measure such a small volume without a very accurate and precise instrument such as a micropipette.

Apparatus	Accuracy
beaker	• used to estimate the volume of a liquid, e.g. approximately 100 cm^3
measuring cylinder	• more accurate than a beaker • measures up to the nearest cm^3, e.g. 99 cm^3
burette	• accurately measures out the volume of a liquid to the nearest ±0.1 cm^3 • scale marked (graduated) in 0.1 cm^3 divisions
pipette	• accurately measures out fixed volumes of liquids, e.g. 20.0 cm^3 or 25.0 cm^3

Table 2.1 Apparatus used to measure volume

• Used to deliver different volumes of liquids, e.g. 24.0 cm^3 or 38.9 cm^3.

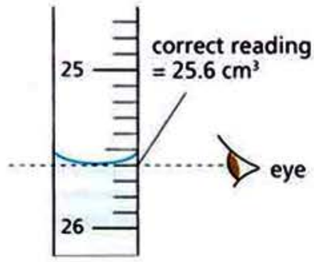


Fig. 2.5 Reading a concave meniscus in a burette

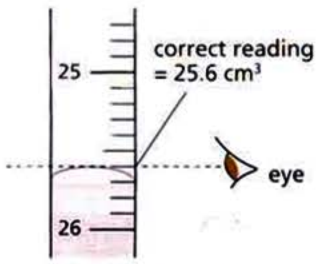


Fig. 2.6 Reading a convex meniscus in a burette

Quick Check

1. The burette shown has a capacity of 50 cm^3. What is the volume of the liquid in the burette?

2. Name the apparatus you would use to measure
   a) 25.8 cm^3 of water.
   b) 25.0 cm^3 of water.

How should we read the volume of a liquid?

When water or a solution is placed in a glass container, it forms a curved surface called a **meniscus**. A meniscus may be concave or convex.

To read the volume of a liquid, align your eyes to the liquid level. If the meniscus is concave, read off the scale at the bottom of the meniscus (Fig. 2.5). If the meniscus is convex, read off the scale at the top of the meniscus instead (Fig. 2.6).

The S.I. unit for volume is the **cubic metre (m^3)**. The **cubic centimetre (cm^3)**, the **litre (l)** and the **millilitre (ml)** are also used.

- 1 litre (l) = 1000 millilitres (ml)
- 1 litre (l) = 1 cubic decimetre (dm^3)
  = 1000 cubic centimetres (cm^3)
- 1 cubic metre (m^3) = 1000 litres (l)

2.5 | Collecting Gases and Measuring Volume of Gases

In some experiments, gases are given off. The gases can be collected and identified to help us find out more about the reactions that have occurred. Sometimes, the gases are collected for use in another reaction.

What methods are used for collecting gases?

How we collect a gas depends on the physical properties of the gas, namely

- solubility — how soluble the gas is in water,
- density — how dense the gas is compared to air.

Table 2.2 shows the solubility and density of some common gases.

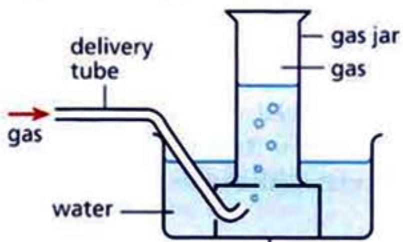
Gas	Solubility in water	Density compared to air
ammonia	extremely soluble	less dense
carbon dioxide	slightly soluble	denser
chlorine	soluble	denser
hydrogen	not soluble	less dense
hydrogen chloride	very soluble	denser
oxygen	very slightly soluble	slightly denser
sulphur dioxide	very soluble	denser

Table 2.2 Solubility and density of common gases

Chem-Aid

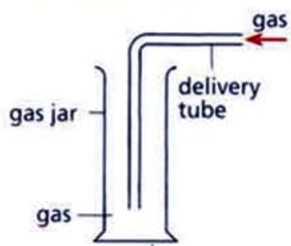
1. A denser gas has heavier gas particles than a gas that is less dense.
2. As a general guide, a gas is less dense than air if its molecular mass is less than 30.

Figs. 2.7, 2.8 and 2.9 show three methods of collecting gases.



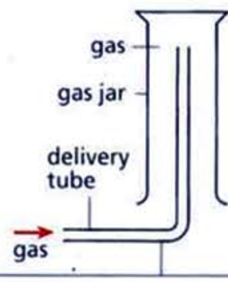
Displacement of water is suitable for collecting *gases that are insoluble or slightly soluble in water.* Carbon dioxide, hydrogen or oxygen can be collected by this method.

Fig. 2.7 Displacement of water



Downward delivery should be used to collect *gases that are soluble in water and denser than air*, such as chlorine and hydrogen chloride.

Fig. 2.8 Displacement of air — downward delivery

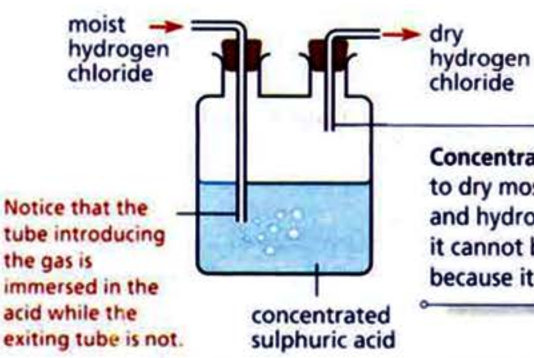


Upward delivery is used to collect *gases that are soluble in water and less dense than air*, such as ammonia.

Fig. 2.9 Displacement of air — upward delivery

How do we collect a dry sample of a gas?

Sometimes, we need to use a dry gas in an experiment. We can dry a gas by passing it through a **drying agent**. Some commonly used drying agents are **concentrated sulphuric acid**, **quicklime (calcium oxide)** and **fused calcium chloride**. Figs. 2.10, 2.11 and 2.12 show three sets of apparatus and the chemicals used to dry gases.



Concentrated sulphuric acid is used to dry most gases including chlorine and hydrogen chloride. However, it cannot be used to dry ammonia because it reacts with ammonia.

Fig. 2.10 Drying a gas with concentrated sulphuric acid

Concentrated sulphuric acid is corrosive to the skin, eyes and clothing. Wear safety goggles and gloves if you are using concentrated sulphuric acid for experiments. It is best to conduct these experiments in a fume cupboard.

Quick Check

Ammonia, chlorine and hydrogen chloride cannot be collected by displacement of water. Why?

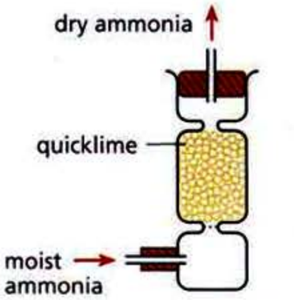


Fig. 2.11 Quicklime (calcium oxide) is used to dry ammonia.

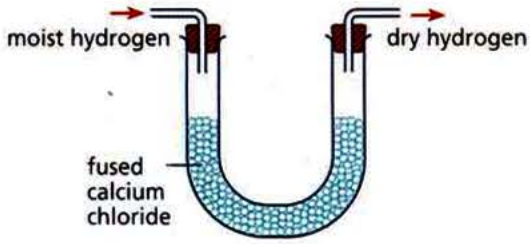


Fig. 2.12 Fused calcium chloride (i.e. calcium chloride that has been previously heated) can be used to dry most gases.

Link

For a reaction that produces a gas, how can you measure the speed at which the reaction is occurring? Find out in chapter 18.

How do we measure the volume of a gas?

It is possible to find out how far a reaction has proceeded by recording the volume of the gas produced in a reaction.

A **gas syringe** is used to measure the volume of a gas. The gas syringe measures a maximum volume of 100 cm^3. It is made up of two parts, the barrel and the plunger.

At the start of an experiment, the plunger is pushed in fully to expel any gas in the syringe (Fig. 2.13). As the gas from an external source enters the syringe, it pushes the plunger outwards (Fig. 2.14).

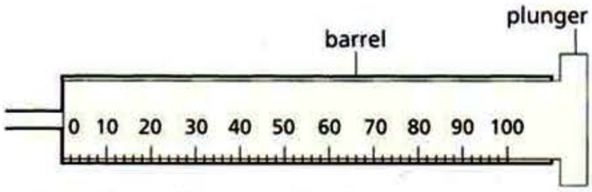


Fig. 2.13 Gas syringe at the start of an experiment

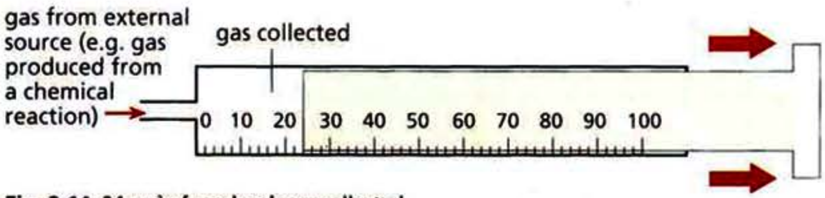


Fig. 2.14 24 cm^3 of gas has been collected.

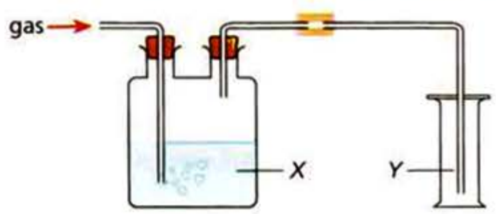
Key ideas

1. A stopwatch or stopclock is used to measure time accurately.
2. A thermometer is used to measure temperature accurately.
3. An electronic balance is used to measure mass accurately.
4. A burette or a pipette is used to measure the volume of a liquid accurately.
5. A gas syringe is used to measure the volume of a gas.
6. The method of collecting a gas depends on the density and solubility of the gas.

Test Yourself 2.1

Worked Example

The apparatus shown was used to collect a dry gas. What are *X* and *Y*?



	X	*Y*
A	concentrated sulphuric acid	ammonia
B	concentrated sulphuric acid	carbon dioxide
C	water	carbon dioxide
D	water	ammonia

Thought Process

A is incorrect as ammonia reacts with concentrated sulphuric acid.
C and D are incorrect because water would dissolve ammonia and make carbon dioxide wet. Carbon dioxide is denser than air and is collected by downward delivery. It does not react with concentrated sulphuric acid. Thus, B is the answer.

Answer

B

Question

Methane is insoluble in water and is less dense than air. Describe, with the aid of a diagram, how you would collect a dry sample of methane gas.

Concept Map

Quantities commonly measured in chemistry experiments

Time
- in seconds (s), minutes (min) or hours (h)
- using stopwatches or stopclocks

Mass
- in grams (g) or kilograms (kg)
- using beam balances and electronic balances

Temperature
- in degree Celsius (°C)
- using thermometers or temperature sensors

Volume
- in cubic centimetres (cm^3)
- using beakers, measuring cylinders, burettes, pipettes and syringes

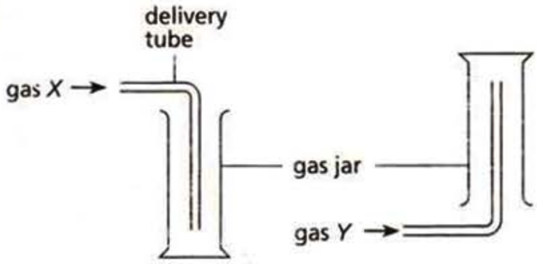
Exercise 2

Foundation

1. What would a chemist use to measure exactly 25.5 cm^3 of dilute hydrochloric acid?

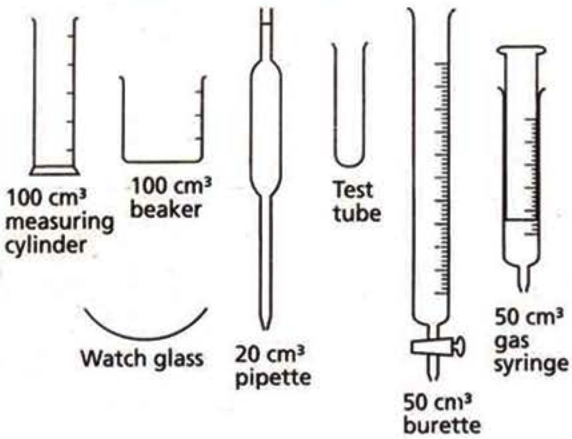
   A Burette
   B Beaker
   C Measuring cylinder
   D Pipette

2. Gas *X* and gas *Y* can be collected using the apparatus shown below. What can you infer about *X*?



   A *X* is less dense than air.
   B *X* is slightly soluble in water.
   C *X* is denser than *Y*.
   D *X* is less soluble in water than *Y*.

3. From the common laboratory apparatus shown below, select the best apparatus to perform each of the following tasks. (You may use any of the apparatus once, more than once or not at all.)



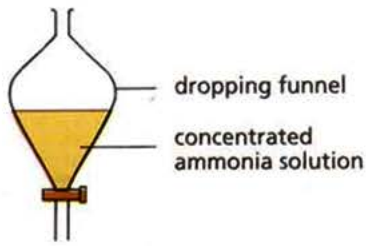
   a) Measuring 80 cm^3 of alcohol
   b) Measuring 35 cm^3 of hydrogen
   c) Measuring 21.1 cm^3 of vinegar
   d) Boiling 40 cm^3 of a non-flammable liquid
   e) Holding 2 g of a solid for weighing
   f) Holding 5 cm^3 of water for heating

Challenge

1. Imagine you are working in a laboratory that makes tablets for relieving indigestion. The tablets work by reacting with water or acid in the stomach. Carbon dioxide gas is produced as a result. You are asked to determine if taking the tablets with alcohol will affect the speed at which the tablets work.

   a) State the aim of your investigation.
   b) Describe how you will carry out your investigation. Include diagrams of your experimental set-up. You can use any or all of the apparatus below.
      - 3 (100 cm^3) conical flasks
      - 3 balloons
      - 3 rubber bands
      - 3 indigestion tablets
      - water
      - alcohol
      - measuring cylinder
      - stopwatch
      - mass balance

2. Ammonia gas can be made by adding concentrated ammonia solution from a dropping funnel, shown below, to sodium hydroxide pellets placed in a round-bottomed flask.

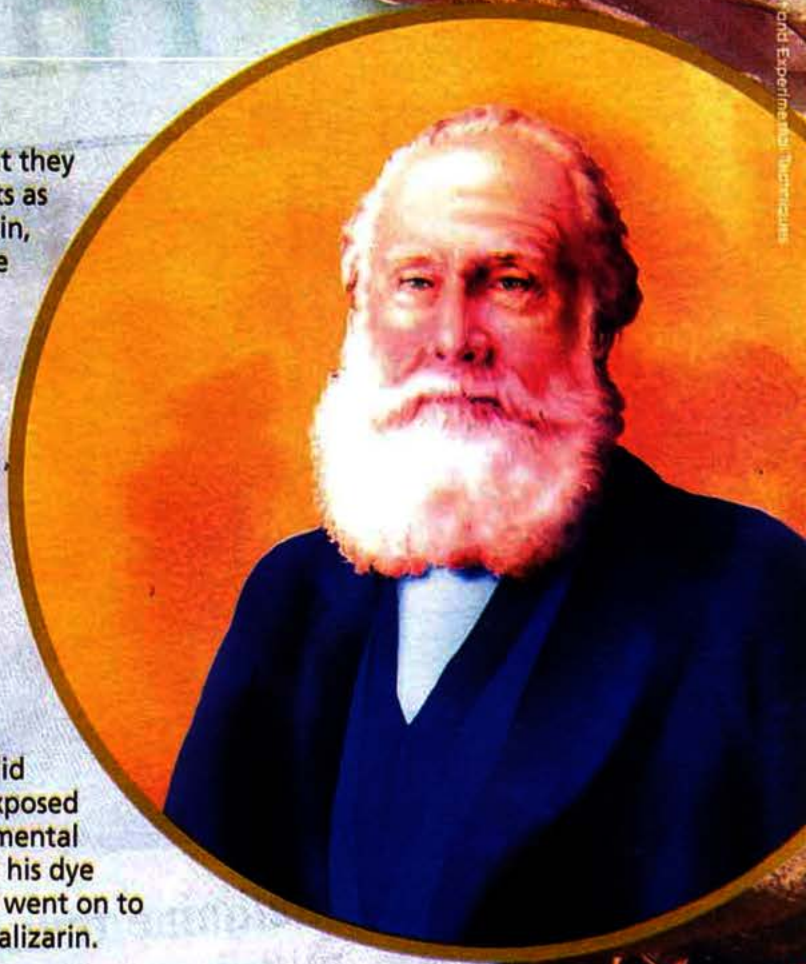


   The ammonia is dried by passing it through solid calcium oxide (quicklime). Ammonia is collected in a gas syringe.

   a) i) Copy the diagram shown above and complete it to show how ammonia can be prepared by this method.
      ii) Suggest why the volume of gas collected initially is not pure ammonia.
   b) Ammonia is less dense than air and very soluble in water. Draw a labelled diagram to show how ammonia can be collected by displacement of air. (Only the collection apparatus is required.)

Chemistry Today

The mark of outstanding scientists is that they do not consider unsuccessful experiments as failures. One such scientist, William Perkin, made an important discovery because he examined his failed experiment from other angles.



In 1856, William Perkin, was asked to see if he could make the anti-malarial drug, quinine, from aniline, a waste product of coal tar. Perkin failed to make quinine but he obtained a black powder. In trying to discover what this powder was, he dissolved it in alcohol. He obtained a purple coloured solution and wondered if it could be used as a dye. Perkin tried and found that the solution coloured silk with a beautiful mauve colour. The dye was also fast (it did not wash out with soap or fade when exposed to sunlight)! Using his accidental experimental results, Perkin built a factory to produce his dye and made a vast fortune for himself. He went on to discover other dyes such as the red dye, alizarin.

It was later found that these dyes could be used to stain bacteria which made bacteria easier to see under the microscope. Tuberculosis and cholera bacilli were discovered using this technique.

CRITICAL THINKING

1. Imagine you were the chemist whom Perkin put in charge of his first dye factory. You have to ensure that a consistent purple dye is always produced from aniline. Perkin asks you to list the quantities that would have to be kept constant during the production process. Make your list.

2. Why do you think synthetic dyes have mostly replaced natural substances as dyes nowadays?

Chapter **3**

Purification and Separation


THE TIMES

WEDNESDAY SEPTEMBER 30, 2005

Medicine found to contain banned impurities withdrawn from shelves islandwide

Chapter Outline

- 3.1 Determining Purity
- 3.2 Chromatography
- 3.3 Separation Techniques
- 3.4 Separating a Solid from a Liquid
- 3.5 Separating Solids
- 3.6 Separating a Liquid from a Solution
- 3.7 Separating Liquids

The Health Sciences Authority (HSA) of Singapore is the government body that regulates food and drug safety in Singapore. Samples of food and medicines are regularly tested in HSA laboratories to make sure that they do not contain banned impurities and are safe for use.

What is a pure substance? How can we test the purity of a substance? How can we make a substance pure? You will find out in this chapter.

3.1 | Determining Purity

Every day, you will come across different substances. Some of them are pure. Others are not pure — they are mixtures.

How is a pure substance different from a mixture?
A **pure substance** is *made up of only one substance* and is not mixed with any other substance. For example, white diamond is made only of carbon.

White diamonds are pure carbon but coloured diamonds are impure because they contain substances other than carbon.

Why doesn't a chemist consider 'pure' orange juice to be a pure substance?

Link
Learn more about mixtures in chapter 4.

Chem-Aid
Impurities are the materials that need to be removed in order to make a substance pure.

In nature, many substances are not pure. The air we breathe is not pure. Air is a mixture of gases such as carbon dioxide, oxygen, nitrogen and water vapour. A **mixture** is a *substance that contains two or more substances that are not chemically combined.*

Is it important to determine whether a substance is pure?
It is important to determine the purity of substances. For example, impurities in medicine must be detected as they may produce undesirable effects when the medicine is consumed.

Purity is also important in the food and beverage industry. Chemicals such as preservatives and dyes are added into food and beverages to make them last longer, taste better or simply look more attractive. It is important to ensure that our food and drinks contain only chemicals that are safe for consumption. Hence, it is important to ensure the purity of these chemicals.

How do we determine whether a substance is pure?
Modern laboratories use many elaborate instruments to determine if a substance is pure. In your school laboratory, you can test if a substance is pure by checking its melting point and boiling point, or by using chromatography. Chromatography is an experimental technique that will be described in the next section.

Pure Solids have Fixed Melting Points

A solid is pure if it has an exact and constant (or fixed) melting point. This means that *a pure solid will melt completely at one temperature.*

Permitted dyes and flavourings give this beverage a unique colour and taste.

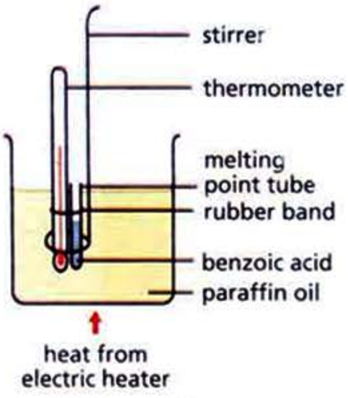


Fig. 3.1 Determining the purity of benzoic acid by checking its melting point

Take benzoic acid for example. Benzoic acid is found naturally in plums, cinnamon and apples. It can kill bacteria and is artificially produced for use as a preservative in cosmetics and food. Some products such as toothpaste and heat-sterilised food need extremely pure benzoic acid as a preservative. The melting point of pure benzoic acid is 122 °C. The apparatus shown in Fig. 3.1 can be used to determine if a sample of benzoic acid is pure.

What is the effect of impurities on melting point?

Impurities affect the melting point of a substance in two ways.

- They lower the melting point. The greater the amount of impurities, the lower the melting point of the substance.
- They cause melting to take place over a range of temperatures. For example, an impure sample of benzoic acid may melt over a temperature range of 118 – 121 °C.

ScienceSkills

1. Examine the set-up in Fig. 3.1. Why do you think paraffin oil is used instead of water?
2. Why is it important to stir the paraffin oil during the experiment?

Pure Liquids have Fixed Boiling Points

A liquid is pure if it has an exact and constant (fixed) boiling point. Fig. 3.2 shows the apparatus for determining whether liquid propanone is pure. The boiling point of pure propanone is 56 °C.

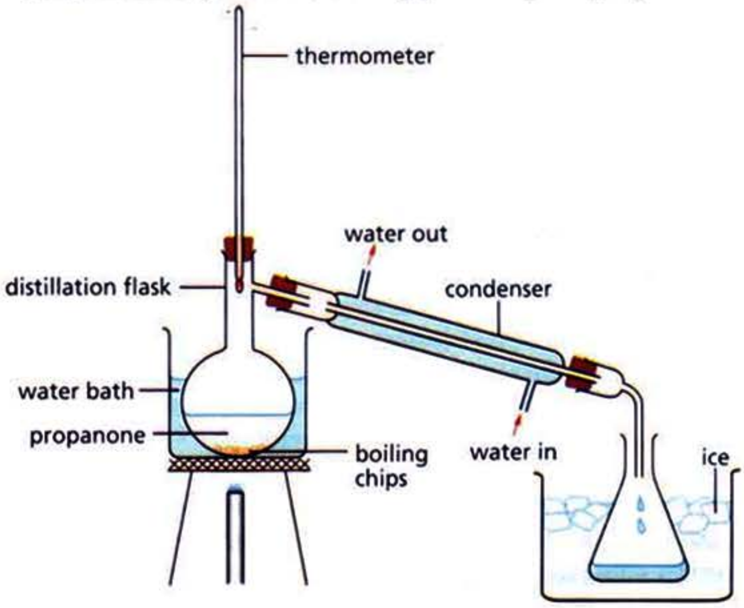


Fig. 3.2 Determining the boiling point of propanone

Chem-Aid

Propanone is highly flammable. Avoid heating flammable liquids directly with a Bunsen flame. Use a water bath instead.

ScienceSkills

Ethylbenzene has a boiling point of 136 °C. Explain why its boiling point cannot be determined using the apparatus shown in Fig. 3.2.

What is the effect of impurities on boiling point?

If a liquid is impure, its boiling point will increase. The liquid will also boil over a range of temperatures. The greater the amount of impurities, the higher the boiling point of the liquid.

What is the effect of pressure on boiling point?

If the pressure acting on a liquid is increased, the boiling point of the liquid is raised. On the other hand, if the pressure is decreased, the boiling point is lowered.

Key ideas

1. The purity of a substance can be determined by measuring its melting point or its boiling point.
2. A pure solid melts completely at a fixed temperature.
3. A pure liquid boils at a fixed temperature.
4. Impurities lower the melting point of a solid and increase the boiling point of a liquid.
5. When pressure increases, the boiling point of a liquid increases. When pressure decreases, the boiling point of a liquid decreases.

Test Yourself 3.1

Worked Example

An impure sample of *X* melts at around 90 °C. What is most likely to be the melting point of *X*?

A Between 90 °C and 110 °C.
B Between 80 °C and 90 °C.
C Below 80 °C.
D Cannot be determined.

Thought Ptocess

Impurities decrease the melting point of a pure substance. Hence, the melting point of *X* should be above 90 °C.

Answer

A

Questions

1. The following data is obtained experimentally:

Acid	Melting point (°C)	Boiling point (°C)
methanoic acid	8	101
ethanoic acid	17	118

An unknown substance, *X*, melts at around 15 °C and boils at around 121 °C. Based on the data given above, suggest what *X* might be. Briefly explain your answer.

2. An unidentified substance, *Y*, was thought to be one of the following three compounds.

Compound	*A*	*B*	*C*
Melting point (°C)	132	133	134

Y was mixed with each of the above compounds and the melting point of each mixture was determined. The following results were obtained.

Mixture	*Y* + *A*	*Y* + *B*	*Y* + *C*
Melting point (°C)	114 – 129	132 – 133	86 – 99

Deduce the identity of *Y*. Briefly explain your answer.

3.2 | Chromatography

Food colouring is often used to colour the things that we eat and drink. Food colouring is a mixture of coloured compounds called dyes. The government ensures that only dyes that are safe for us to consume are used in food colouring. How can the government test whether all the dyes in food colouring are safe for us to consume? To do so, they can make use of a technique called **chromatography**. Even you can perform chromatography!

Chem-Aid

ScienceSkills

Ink is a mixture of different dyes.

1. A small drop of ink is placed in the centre of a piece of filter paper.
2. When the drop has dried up, another drop is added at exactly the same spot and is allowed to dry.
3. Ethanol, which is a solvent, is slowly added drop by drop onto the spot (Fig. 3.3).
4. The addition of ethanol causes the spot of ink to slowly spread out into different coloured rings (Fig. 3.4). Each ring represents a different dye that made up the ink. You have just separated the dyes in the ink!

The technique of using a solvent to separate a mixture into its components is called **chromatography**. The technique that you have just used is called **paper chromatography**.

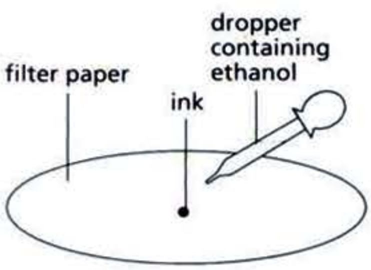


Fig. 3.3 A simple chromatography experiment

Paper chromatography can be used to separate dyes in ink, pigments in plants, amino acids obtained from proteins, to identify poisons (e.g. pesticides) or drugs, and to detect traces of banned substances in food.

There is another way that paper chromatography is used to separate the dyes in ink. This method allows the solvent to ascend or travel up the chromatography paper.

How can we use chromatography to separate dyes in food colouring?

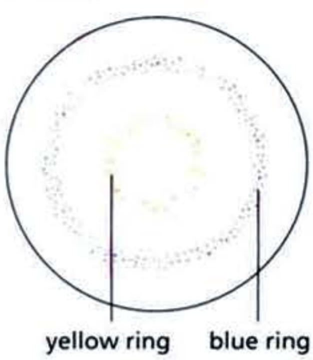


Fig. 3.4 The separation of ink by chromatography

Experiment 1

To separate the dyes in green food colouring using ascending paper chromatography, we set up the apparatus as shown in Fig. 3.5.

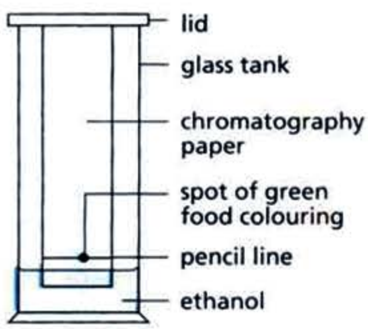


Fig. 3.5 Ascending paper chromatography

Dyes are also used to colour cloth.

How does chromatography work?

- A spot of food colouring is applied to the chromatography paper.
- Once the chromatography paper is dipped in ethanol (the solvent), it soaks up ethanol.
- Ethanol that is soaked up by the paper dissolves the dyes. It continues to travel up the paper, carrying the dyes along.
- A dye that is not very soluble in ethanol will not be carried far along the paper.
- A dye that is very soluble in ethanol will be carried far along the paper.
- Coloured spots are left in different places on the paper at the end of the experiment.

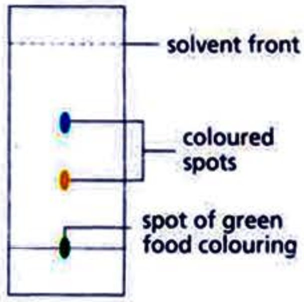


Fig. 3.6 A chromatogram of separated dyes in green food colouring

How do we interpret the result of chromatography?

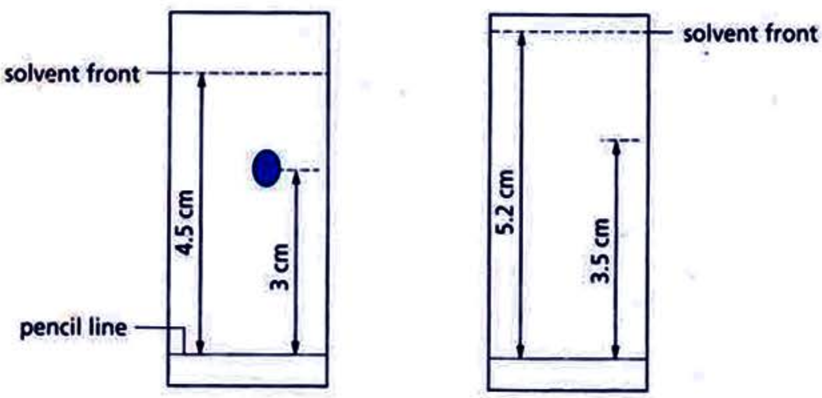
Fig 3.6 shows a typical result when the dyes in food colouring are separated by chromatography. *The chromatography paper with the separated components* is called a **chromatogram**. The chromatogram shows that the food colouring used in the experiment was a mixture of two dyes.

However, if there is only one spot on the chromatogram, it means the substance is pure (Fig. 3.7), that is, it is made up of only one dye.

Fig. 3.7 Chromatogram of a pure substance

What are R_f values?

Look at Fig. 3.8. It shows the chromatograms of the same substance but in Fig. 3.8(b), chromatography was allowed to proceed for a longer time.



$R_f = \frac{3\text{ cm}}{4.5\text{ cm}} = 0.67$

(a) After a short time

$R_f = \frac{3.5\text{ cm}}{5.2\text{ cm}} = 0.67$

(b) After a longer time

Fig. 3.8 Measurement of R_f values

The positions of the solvent front and spot on a chromatogram depend on how long the experiment was allowed to run. However, *the ratio between the distance travelled by the substance and the distance travelled by the solvent is a constant.* This ratio is called the **R_f value** of the substance.

$$R_f = \frac{\text{distance travelled by the substance}}{\text{distance travelled by solvent}}$$

The R_f value of a substance does not change as long as chromatography is carried out under the same conditions (i.e. same solvent and same temperature). This property allows us to easily identify a substance on a chromatogram. Why do you think different substances have different R_f values?

Using Paper Chromatography to Analyse a Sample

In analysing a sample, we identify the components present in the sample. For example, as mentioned previously, food colouring is used to colour food and drinks. The government needs to check whether all of the dyes in the food colouring are safe for human consumption.

The following steps are taken to analyse a sample of food colouring:

- Paper chromatography is used to separate the dyes in the sample.
- Each dye can then be identified by comparing its position in the chromatogram with that of a known dye.
- Each dye can also be identified by comparing its R_f value with the R_f value of a known dye.
- Chemists can then check whether the dyes are permitted for use in food.

Fig. 3.9(a) Put a drop of each dye onto the chromatography paper.

How can a banned substance present in food colouring be identified?

Fig. 3.9(a) shows a sample of food colouring (labelled '*X*') as well as four dyes (labelled 'A', 'B', 'C' and 'D') on a large sheet of chromatography paper. The four dyes are banned because they are not safe to be consumed. We need to determine whether the food colouring contains any of these banned substances. If it does, it is not safe to be consumed.

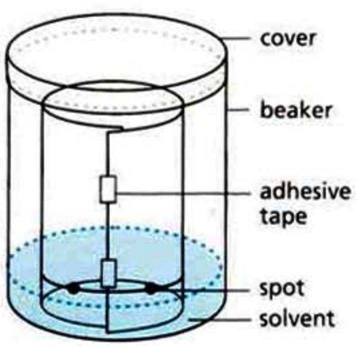


Fig. 3.9(b) Dip the chromatography paper into the solvent.

The sheet of chromatography paper is coiled into a cylinder and secured with adhesive tape. It is put in a beaker containing a suitable solvent (Fig. 3.9(b)). The chromatogram shown in Fig. 3.9(c) is obtained.

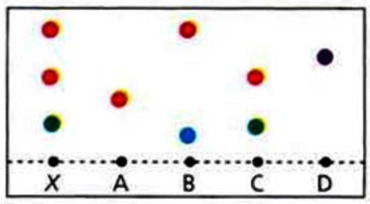


Fig. 3.9(c) The chromatogram produced by the food colouring and dyes

What conclusions can be drawn from this chromatogram?
Dyes A and D are pure. Both dye B and dye C are mixtures of two different dyes. Sample *X* is a mixture of three dyes.

Identical dyes produce spots at the same height and in the same colour on the paper when the same solvent is used (as in this case). Fig. 3.9(c) shows that sample *X* does not contain the banned dyes A, B and D. However, *X* contains the banned dye C. Therefore, it must not be consumed.

How do we perform chromatography on colourless substances?
Chromatography can also be used for colourless substances such as amino acids. To separate and analyse colourless substances, we apply a **locating agent** on a chromatogram.

To separate and analyse amino acids, we follow these steps:
1. Separate the mixture of amino acids by chromatography using a suitable solvent.
2. Before the solvent reaches the top of the paper, stop the chromatography. Dry the paper.
3. Spray a locating agent onto the paper.
4. The locating agent reacts with each of the amino acids to form coloured spots on the paper. By checking the R_f value of each coloured spot, we can identify the different amino acids.

TidBit

Ninhydrin is a chemical used as a locating agent. It gives a blue colouration with amino acids. Colourless substances can also be detected by using ultraviolet radiation.

What are the uses of chromatography?
Chromatography is used to
- separate the components in a sample,
- identify the number of components in a sample,
- identify the components present in a sample,
- determine whether a sample is pure.

Key ideas

1. Chromatography is used to separate or analyse the components in a sample.
2. Chromatography is also used to determine the purity of a sample. A pure sample gives only one spot on a chromatogram.
3. R_f value $= \dfrac{\text{distance travelled by the substance}}{\text{distance travelled by solvent}}$
4. A locating agent is used on a chromatogram to help us see colourless substances.
5. To determine the purity of a substance, we can
   - check its melting point or boiling point,
   - perform chromatography.

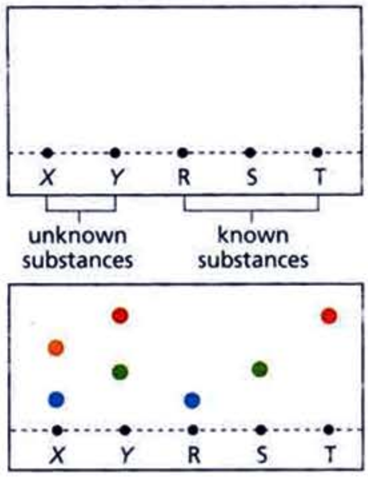
Test Yourself 3.2

Worked Example

A chemist was asked to identify two unknown substances, *X* and *Y*. The chromatograms he made of the mixtures are shown below.

Which statement is true?

A *X* contains R and S.
B *X* contains S and T.
C *Y* contains R and S.
D *Y* contains S and T.



Thought Process

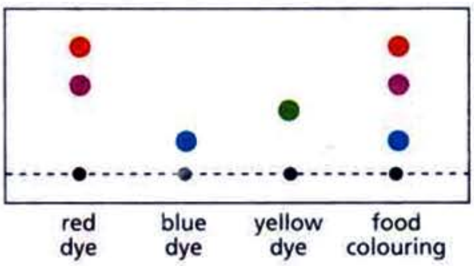
Y contains S and T. These two dyes in *Y* are in the same colour and at the same position on the chromatogram as S and T. *X* contains R and an unknown dye.

Answer

D

Question

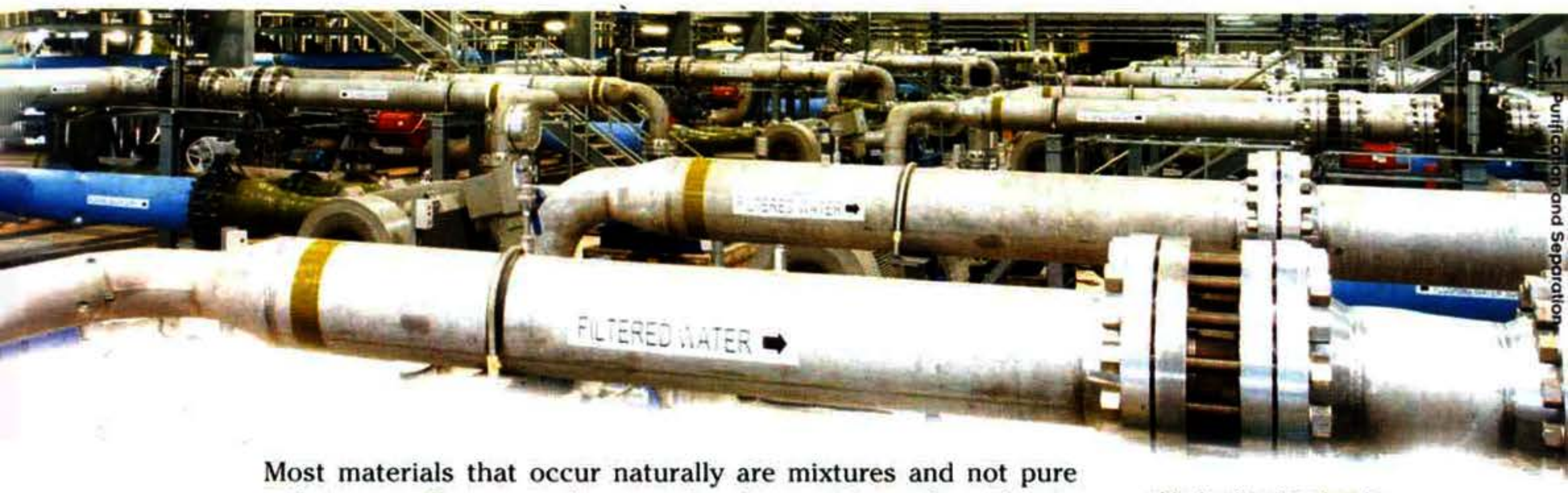
The figure below shows the chromatogram produced by three dyes and a food colouring. What colour would you expect the food colouring to be?



3.3 | Separation Techniques

Chromatography can be used to separate a mixture, identify its components and determine the purity of a substance. It does not allow us to remove or reuse the components of the mixture. However, there are times when we may want to remove or reuse the components of a mixture after separating them.

There are other methods to separate mixtures. These methods allow us to reuse the separated components. Such methods can also be used to separate a mixture into pure substances (purify the mixture). These methods are called **methods of purification** or **separation techniques**.



This desalination plant in Singapore removes salt from seawater.

Most materials that occur naturally are mixtures and not pure substances. For example, seawater is a mixture. In order to produce pure water (which can be consumed) from seawater, we need to remove impurities like salt.

There are several purification techniques we can use to separate mixtures. To decide which technique to use, we need to consider the properties of each substance in the mixture. In the following sections, you will learn how to separate the following:

a) A solid from a liquid
b) Solids
c) A liquid from a solution
d) Liquids

Try it Out

Desalination is an industrial process of removing salt from seawater. Find out the number of desalination plants operating in Singapore. Which separation and purification techniques are used during desalination?

3.4 | Separating a Solid from a Liquid

Look at the two pictures on the right. If you need to remove the pebbles and the sand from water, would you use the same technique for both?

Pebbles in water

Fine sand in water

Decanting

The simplest way to separate the water from the pebbles is to just pour the water away. This method is called **decanting** (Fig. 3.10). We use decanting to separate a dense, insoluble solid from a liquid. Decanting is carried out every day during many activities. For example during cooking, water is poured off various cooked foods by decanting.

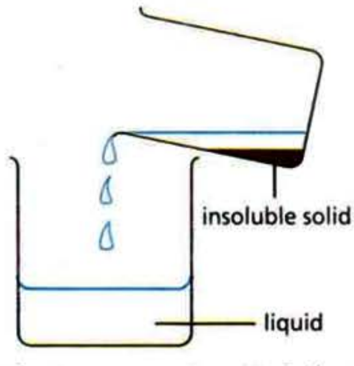


Fig. 3.10 Decanting to separate a dense, insoluble solid from a liquid

Filtration

To separate sand from water, we use a method known as **filtration**. In general, we use filtration to separate small solid particles from a liquid. Examples of small solid particles include sand, clay, dust particles and **precipitates** (small solid particles produced in a liquid by chemical reactions).

Chem-Aid

1. **Precipitate** refers to insoluble solid particles in a liquid.
2. A **suspension** is a mixture where solid particles are found throughout the liquid.

How can we remove a precipitate from a reaction mixture by filtration?

A white precipate of lead(II) sulphate is produced when we add sodium sulphate solution to lead(II) nitrate solution. Decanting the mixture will not be effective as the mixture is a **suspension**. Instead, filtration should be used. Experiment 2 illustrates how filtration is done.

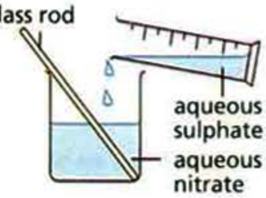
Experiment 2

To separate lead(II) sulphate from a reaction mixture

Procedure

1 Prepare the lead(II) sulphate.

Pour 30 cm^3 of sodium sulphate into a beaker containing 50 cm^3 of lead(II) nitrate solution. Stir the reaction mixture with a glass rod.



2 Fold the filter paper.

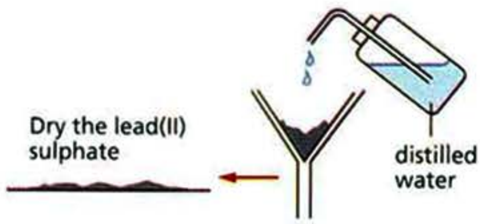
Fold a piece of filter paper as shown. Place it in a filter funnel. Moisten it with a little distilled water.

3 Filter the mixture.

Pour the reaction mixture into the filter funnel that is lined with filter paper. Collect the filtrate that passes through the filter paper in a conical flask.

4 Collect the residue.

The residue is lead(II) sulphate. Wash the residue with distilled water. Then dry the residue on a piece of filter paper.



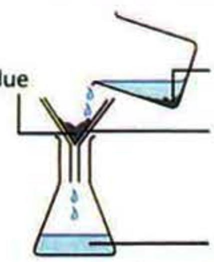
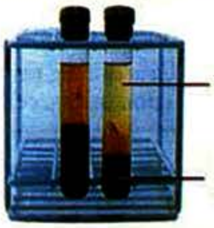


Fig. 3.11 Using filtration to separate lead(II) sulphate from a reaction mixture

TidBit

Sometimes, blood needs to be separated into its components before it can be tested. To do so, we use a machine called a centrifuge.

In the centrifuge, test tubes are rotated, causing solid particles suspended in the liquid to gather at the bottom of the tubes. The solution can be decanted from the tubes without disturbing the solid.



When blood has been centrifuged, it is separated into blood cells and a liquid called plasma.

A solid can be separated from a liquid by filtration because the filter paper acts as a sieve. A liquid can pass through the pores (small holes) of the filter paper but a solid cannot do so (Fig. 3.12).

Upon filtration, *the solid that remains on the filter paper* is called the **residue**. *The liquid or solution that passes through the filter paper* is called the **filtrate**.

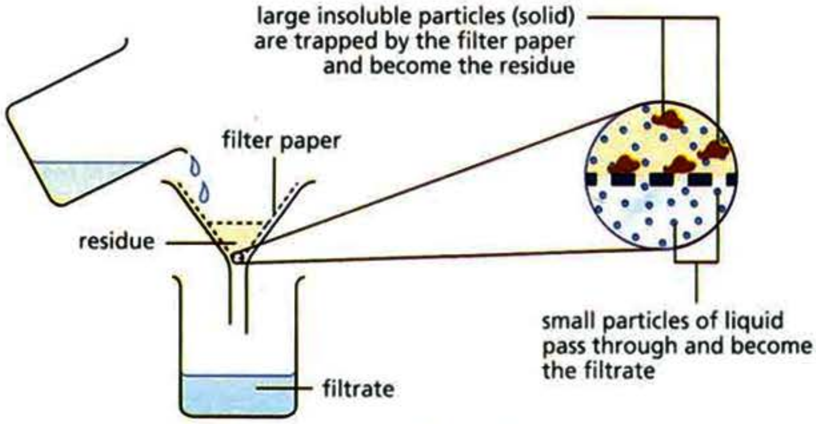


Fig. 3.12 How filtration works

Try it Out

In the production of NEWater, one of the processes used is known as microfiltration. Find out more about microfiltration at http://www.pub.gov.sg/NEWater.

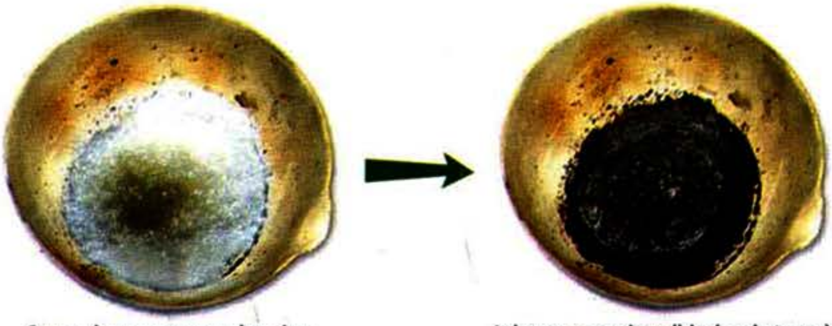
Evaporation to Dryness and Crystallisation

When a solid is insoluble in water, the easiest way to separate it from a liquid is by filtration. However, many substances, like common salt, dissolve in water to form solutions. To separate such substances, we evaporate water from the solution. This separation technique is called **evaporation to dryness**. For example, when we evaporate salt solution to dryness, we recover solid salt. This method can be used to recover salt from seawater.

However, many substances **decompose** when they are heated strongly. For example, sugar will decompose to give water and carbon when it is heated strongly. Most crystals, such as copper(II) sulphate crystals, give off water to become powders when heated. For such substances, evaporation to dryness is not a good method of purification. Also, when all the water is removed during evaporation, any soluble impurities present will be left on the crystals.

Chem-Aid

When a substance decomposes, it breaks down to form simpler substances.



Sugar decomposes on heating. **It becomes carbon (black substance).**

The best method of obtaining a pure solid sample from its solution is **crystallisation**.

How do we purify by crystallisation?

In crystallisation, water is removed by heating the solution. Heating is stopped at the stage when a hot **saturated** solution is formed. If the resulting solution is allowed to cool to room temperature, the dissolved solid will be formed as pure crystals.

Chem-Aid

A saturated solution is a solution that contains the maximum amount of dissolved solute at a given temperature. No more solute can be dissolved in the solution.

How do we test for a saturated solution?

A clean glass rod can be used to test whether a solution is saturated. It is dipped into the solution and removed. There will be a small amount of solution on the rod. If small crystals form on the rod as the solution cools, the solution is saturated. We say the solution is at its **saturation point** or **crystallisation point**.

Quick Check

The easiest way to obtain a solute from a solution is by evaporating all the solvent. Give **two** reasons why chemists often avoid using this method.

The flow chart below (Fig. 3.13) shows the steps involved in obtaining copper(II) sulphate crystals from copper(II) sulphate solution.

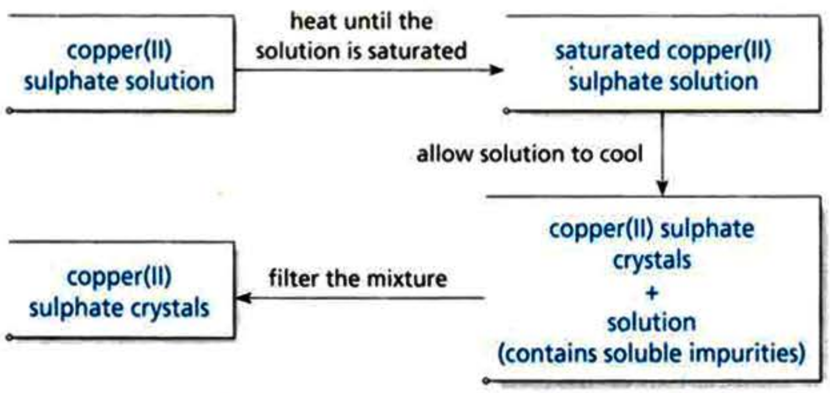


Fig. 3.13 The preparation of copper(II) sulphate crystals

Key ideas

1. Decanting is used to separate a dense, insoluble solid from a liquid.
2. Filtration is used to separate a mixture of a solid and a liquid.
3. Upon filtration, the solid that remains on the filter paper is called the residue. The liquid that passes through the filter paper is called the filtrate.
4. Substances that do not decompose on strong heating can be purified by evaporation to dryness.
5. Crystallisation is used to purify crystals and substances that decompose on strong heating.
6. A saturated solution can be produced by evaporation.

Copper(II) sulphate crystals

3.5 | Separating Solids

To separate mixtures of two or more solids, we make use of what we know about the differences in their properties.

How do we use filtration to separate two solids?

A mixture of two solids can also be separated by filtration if one of them is soluble in a solvent but the other is not. For example, a mixture of salt and sand can be separated by filtration using water as the solvent. Salt is soluble in water but sand is not.

Experiment 3

To separate a mixture of common table salt and sand

Procedure

1. Pour some distilled water into the mixture of common table salt and sand. Stir and warm the mixture.
2. Pour the warm mixture into a filter funnel lined with filter paper. Collect the filtrate in a conical flask (Fig. 3.14).
3. Wash the residue with a little distilled water to remove all the salt solution from it. The residue is sand.
4. Pour the filtrate into an evaporating dish and evaporate the filtrate to dryness (Fig. 3.15). The white solid left in the evaporating dish is salt.

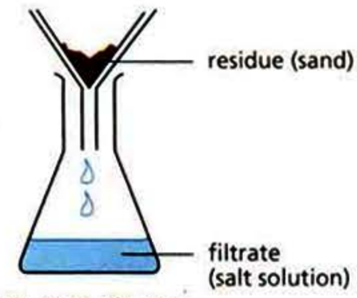


Fig. 3.14 Filtration

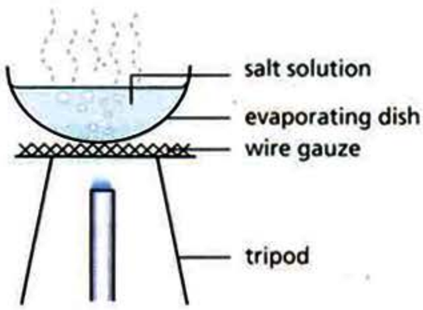


Fig. 3.15 Evaporation

The flow chart below summarises the steps taken to separate salt from sand.

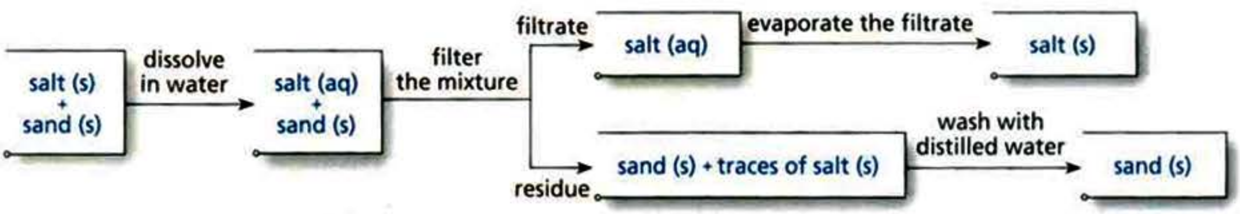


Fig. 3.16 Separating salt from sand

Link

Steel is an alloy made from iron and carbon. Find out more about alloys in chapter 14.

Using a magnet to separate iron from sulphur

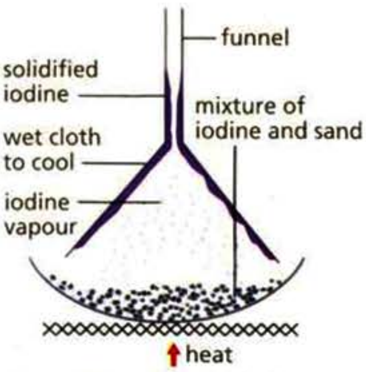


Fig. 3.17 Separating iodine from sand by sublimation

An industrial magnet attracts iron from waste materials.

How do we use a magnet to separate solids?

Some metals are magnetic. The common magnetic metals are iron, nickel and cobalt. Steel is also magnetic. We can use this property to separate these metals and steel from mixtures.

Magnets are used in recycling to recover magnetic materials such as iron and steel from domestic waste. The 'rubbish' is placed on a conveyer belt that passes under a powerful magnet. The magnet attracts these magnetic materials.

How do we use sublimation to separate solids?

Some substances, such as ammonium chloride and iodine, sublime. We can make use of this property to separate a substance that sublimes from one that does not.

For example, to separate a mixture of iodine and sand, we make use of the apparatus shown in Fig. 3.17.

In general, this method is used to separate a substance that sublimes (such as iodine or ammonium chloride) from one with a high melting point (such as sand or sodium chloride).

Key ideas

1. Filtration can be used to separate two solids if only one of them is soluble in a solvent.
2. A magnet can be used to separate magnetic substances, such as iron or steel, from non-magnetic ones.
3. Sublimation is used to separate a substance that sublimes (e.g. iodine) from one with a high melting point (e.g. sand).

Test Yourself 3.3

Worked Example

How would you purify solid sodium nitrate contaminated with solid lead(II) sulphate? Sodium nitrate is soluble in water but lead(II) sulphate is not.

Answer

The flow chart below shows the key steps in the purification process.

A) What do you know about the solids?
- *Lead(II) sulphate is insoluble in water.*
- *Sodium nitrate is soluble in water.*

In the mixture:
Solid sodium nitrate is contaminated with lead(II) sulphate.

B) Can a solvent be used to dissolve one of the solids?

Dissolve in water:
Lead(II) sulphate (s) + sodium nitrate (aq)

C) Which method can separate a liquid from a solid?

Filtration

D) What is the result of filtration?

Filtrate:
Sodium nitrate (aq)

Residue:
Lead(II) sulphate (s)

E) How to get sodium nitrate from solution?

Crystallisation:
- Heat the filtrate until it is saturated.
- Allow the saturated solution to cool.
- Sodium nitrate crystals start to form.

Chem-Aid

To obtain very pure sodium nitrate, the crystals are dissolved in hot water and filtered. The filtrate is then heated to saturation and allowed to cool. The sodium nitrate crystals formed are filtered off and washed with distilled water. The pure crystals are then dried on filter paper. This process is known as **re-crystallisation**.

Questions

1. Describe briefly the steps you would take to obtain samples of
   a) sodium chloride from seawater.
   b) sugar crystals from sugar cane.

2. Potassium nitrate (saltpetre) is used as a fertiliser. It occurs naturally, mixed with soil, in many countries. Potassium nitrate is very soluble in hot water, but not very soluble in cold water. How can we obtain pure potassium nitrate which is free of soil?

3.6 | Separating a Liquid from a Solution

When a solid dissolves in a solvent, a solution is formed. Examples of solutions are salt solution and sugar solution. To collect the solute from the solution, we evaporate the solvent. What method do we use if we need to collect the solvent instead?

How do we obtain a solvent from a solution?

A pure solvent can be separated from a solution by **simple distillation**. Distillation is the *process of boiling a liquid and condensing the vapour.* For example, pure water can be obtained from a salt solution by distillation. The apparatus used for simple distillation in the laboratory is shown in Fig. 3.18.

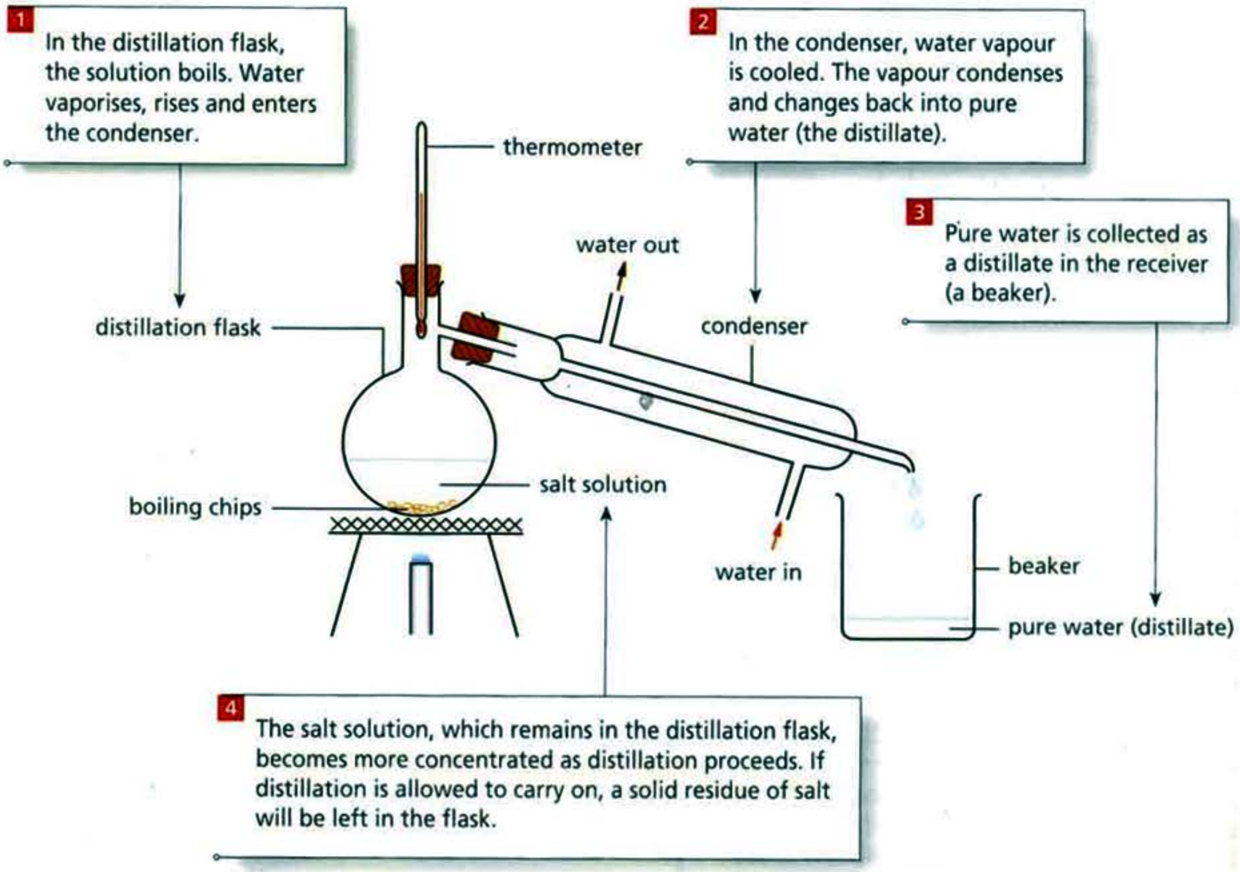


Fig. 3.18 Simple distillation to obtain pure water from salt solution

What steps are taken during distillation?
When setting up the distillation apparatus, note the following procedure:

1. The thermometer should be placed beside the side arm of the distillation flask (Fig. 3.19). It should not be dipped into the solution. This ensures that the thermometer measures the boiling point of the substance that is being distilled.

2. The condenser consists of two tubes: an inner tube and an outer water jacket (Fig. 3.20). Cold running water is allowed to enter the water jacket from the bottom of the condenser and leave from the top.

3. The condenser slopes downwards so that the pure solvent formed can run into the receiver (a beaker).

4. If the distillate is volatile, the receiver can be put in a large container filled with ice. This helps to keep the temperature of the distillate low so that it remains in the liquid state.

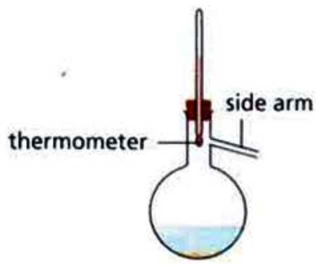


Fig. 3.19 **Distillation flask**

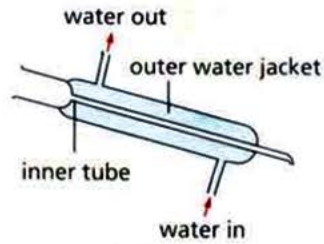


Fig. 3.20 **Condenser**

How does temperature that is recorded vary as a salt solution is distilled?
As the salt solution is heated, its temperature increases (Fig. 3.21). When the solution finally boils, the thermometer records a temperature of 100 °C. This is the temperature of the water vapour. The temperature remains unchanged until all the water has boiled off.

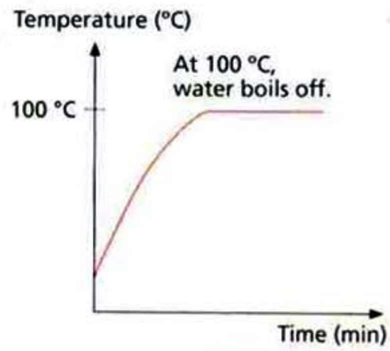


Fig. 3.21 **Graph showing how temperature changes as a salt solution is distilled**

What are the uses of simple distillation?
Simple distillation can be used to:

- recover a solvent from a non-volatile solute. A non-volatile substance, for example, sodium chloride, has a high boiling point.
- separate mixtures of liquids with different boiling points.

However, it is difficult to separate mixtures of liquids whose boiling points differ by less than 20 °C. This problem can be solved if we use fractional distillation to separate the liquids.

3.7 | Separating Liquids

Separating Immiscible Liquids

Liquids that *do not dissolve in each other* are described as **immiscible**. Oil and water are immiscible in each other.

TidBit

When immiscible liquids are shaken, they may appear to mix. In fact, they form an **emulsion**. An emulsion eventually separates into different layers of liquid.

To separate liquids that are not miscible, such as oil and water, we use a separating funnel.

How do we separate immiscible liquids?

To separate immiscible liquids, we use a **separating funnel**. For example, to separate oil and water, we use the following procedure:

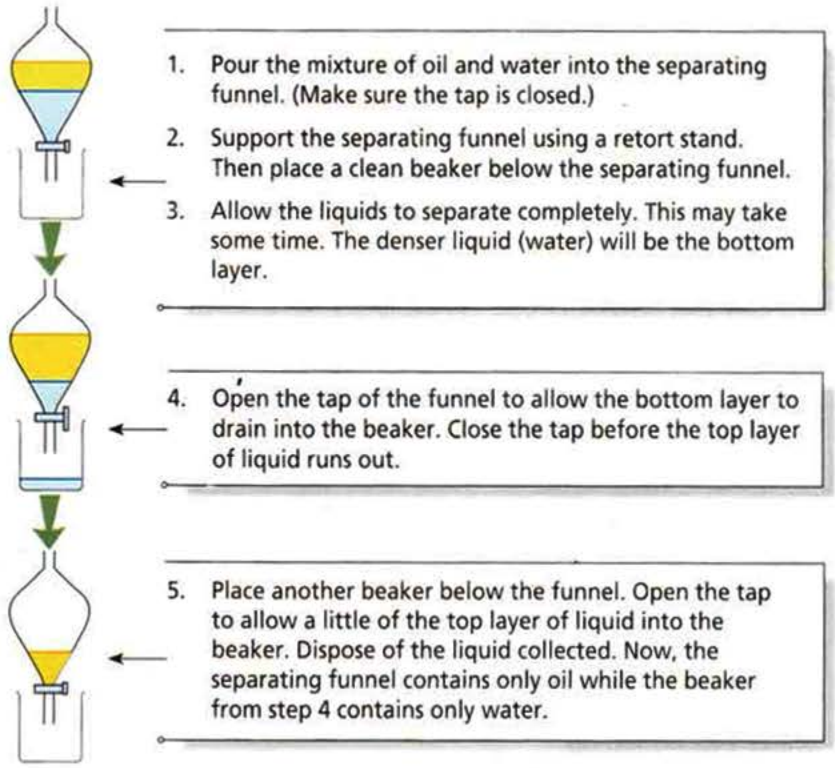


Fig. 3.22 Using a separating funnel to separate oil and water

Separating Miscible Liquids

Ethanol and water *mix together completely to form a solution*. They are said to be **miscible**. If two liquids are miscible, they must be separated by a technique called **fractional distillation**. Fractional distillation can separate a solution of water and ethanol. Fig. 3.23 shows the laboratory apparatus for the fractional distillation of a solution of ethanol and water.

A column, called the **fractionating column**, is attached to the round-bottomed flask and the condenser. Many glass beads in the fractionating column provide a large surface area for vapour to condense on. Other than glass beads, a fractionating column may be filled with plates or a spiral.

During fractional distillation,

- the liquid with the lowest boiling point will distil over to the condenser first,
- the vapours of liquids with higher boiling points condense along the fractionating column and re-enter the round-bottomed flask.

What happens when a solution of ethanol and water undergoes fractional distillation?

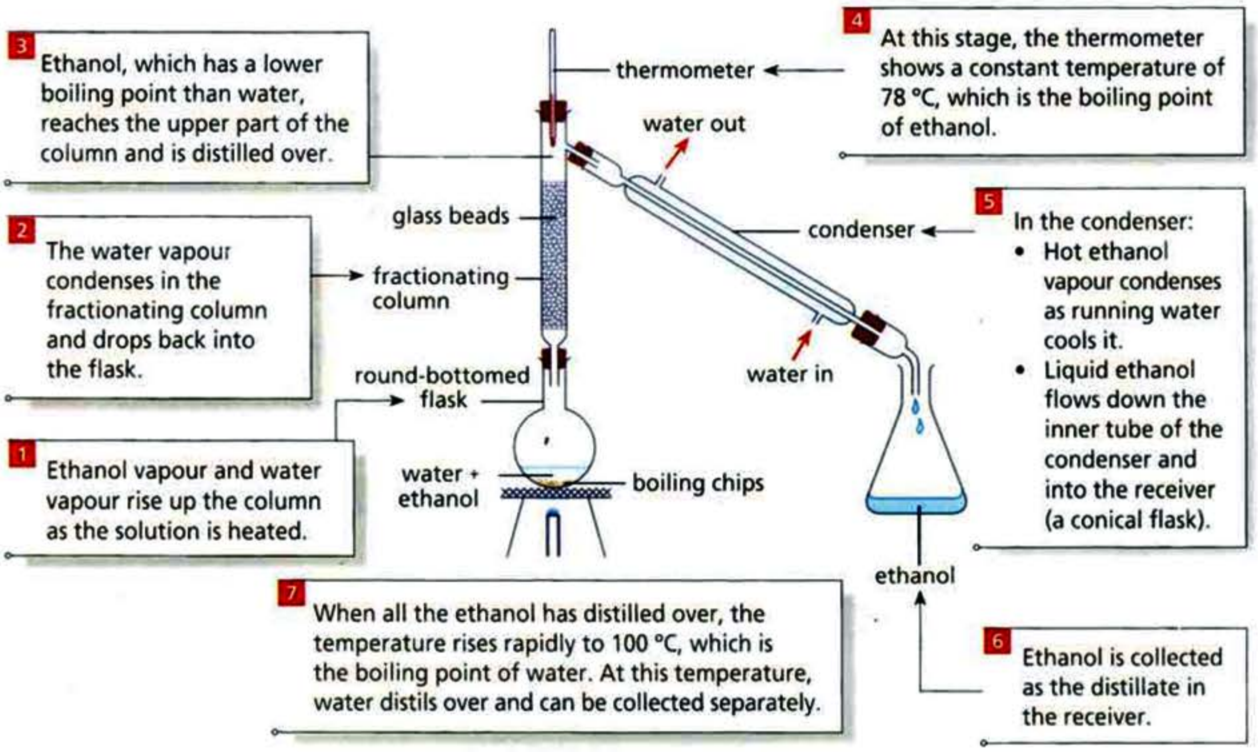


Fig. 3.23 Fractional distillation of a solution of water and ethanol

How does the temperature change as a solution of ethanol and water undergoes fractional distillation?

The temperature of the mixture increases as it is heated. At 78 °C, ethanol distils over. The temperature remains constant until all the ethanol has distilled out of the round-bottomed flask. The temperature then increases until 100 °C. At this temperature, water distils over. The temperature remains unchanged as water is being distilled. Fig. 3.24 shows how the temperature changes as the mixture is heated.

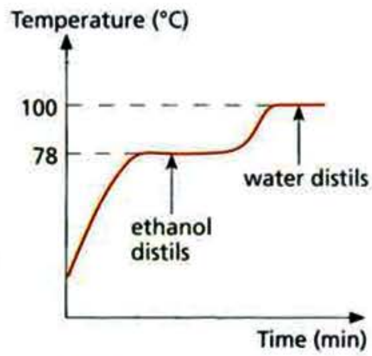


Fig. 3.24 Graph showing how temperature changes as an ethanol/water mixture is fractionally distilled

What are the industrial applications of fractional distillation?

1. Fractional distillation is used in industries to obtain nitrogen, argon and oxygen from air (Fig. 3.25). (See chapter 20).

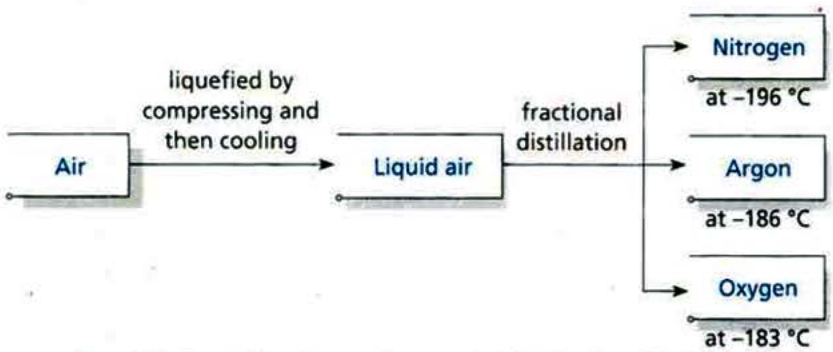


Fig. 3.25 Separating the constituents of air by fractional distillation

Chem-Aid

Fermentation is the process by which microorganisms, such as yeast, change sugars (like glucose) into alcohol (ethanol).

2. Fractional distillation can be used to separate mixtures of liquids such as crude oil. (See chapter 21).

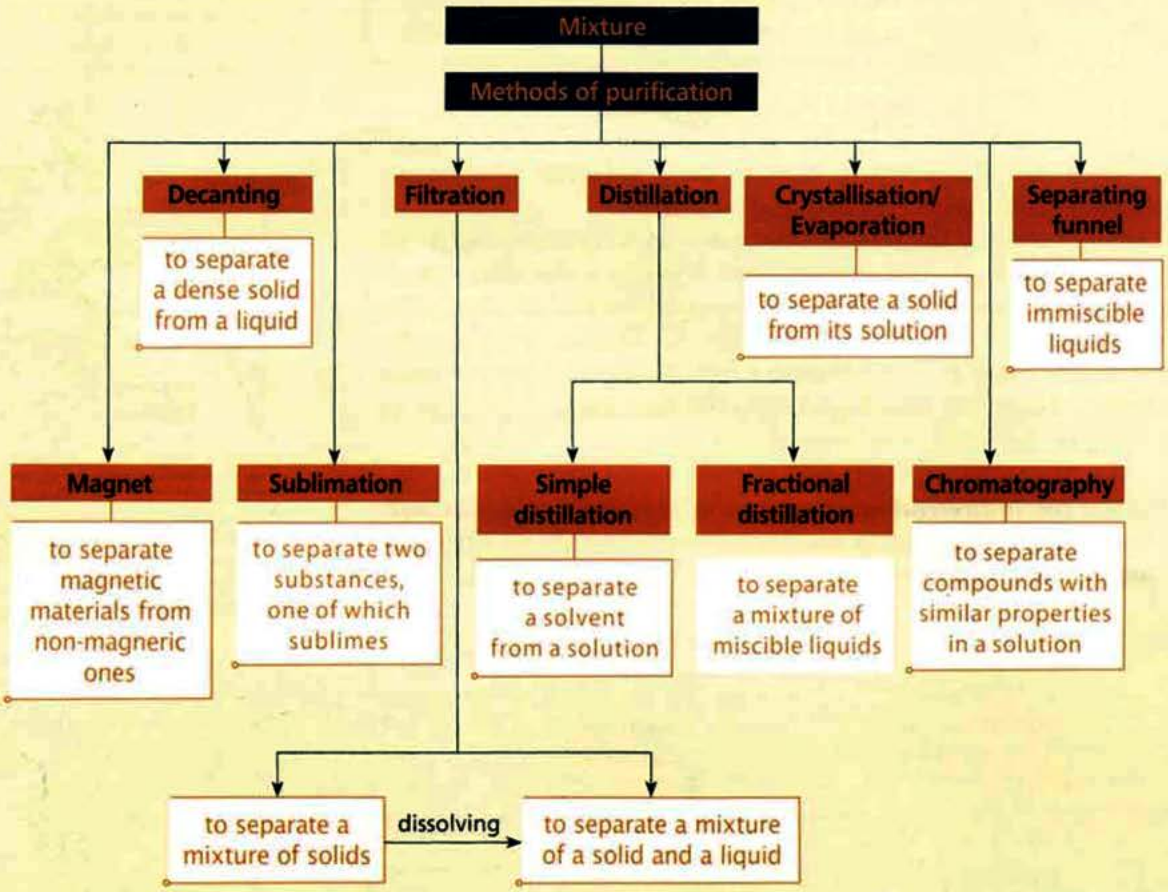
3. Ethanol is formed when glucose solution undergoes fermentation in the presence of yeast. We can separate ethanol from glucose solution by fractional distillation. (See chapter 24).

Key ideas

1. Simple distillation is used to obtain a pure solvent from a solution.
2. Fractional distillation is used to separate a mixture of miscible liquids. The liquid with the lowest boiling point distils over first.
3. During both simple distillation and fractional distillation, vaporisation and condensation occur:

$$\text{Liquid} \xrightarrow[\textit{vaporisation}]{\text{(heating)}} \text{Vapour} \xrightarrow[\textit{condensation}]{\text{(cooling)}} \text{Liquid}$$

4. There are various methods of purification. To choose the most suitable method, you must consider the properties of the components of the mixture.



Test Yourself 3.4

Worked Example

The table below gives information about the various fractions obtained when crude oil is fractionally distilled.

Fraction	Boiling point (°C)
bitumen	over 359
fuel gas	below 40
gasoline	40 – 100
heavy gas oil	300 – 350
kerosene	160 – 250
light gas oil	250 – 300
naphtha	75 – 150

a) Which fraction would be distilled over first? Why?
b) Which fraction will be collected at the bottom of the column? Why?
c) What two other pieces of information do you need to know in order to decide whether these fractions are solids, liquids or gases at room temperature?

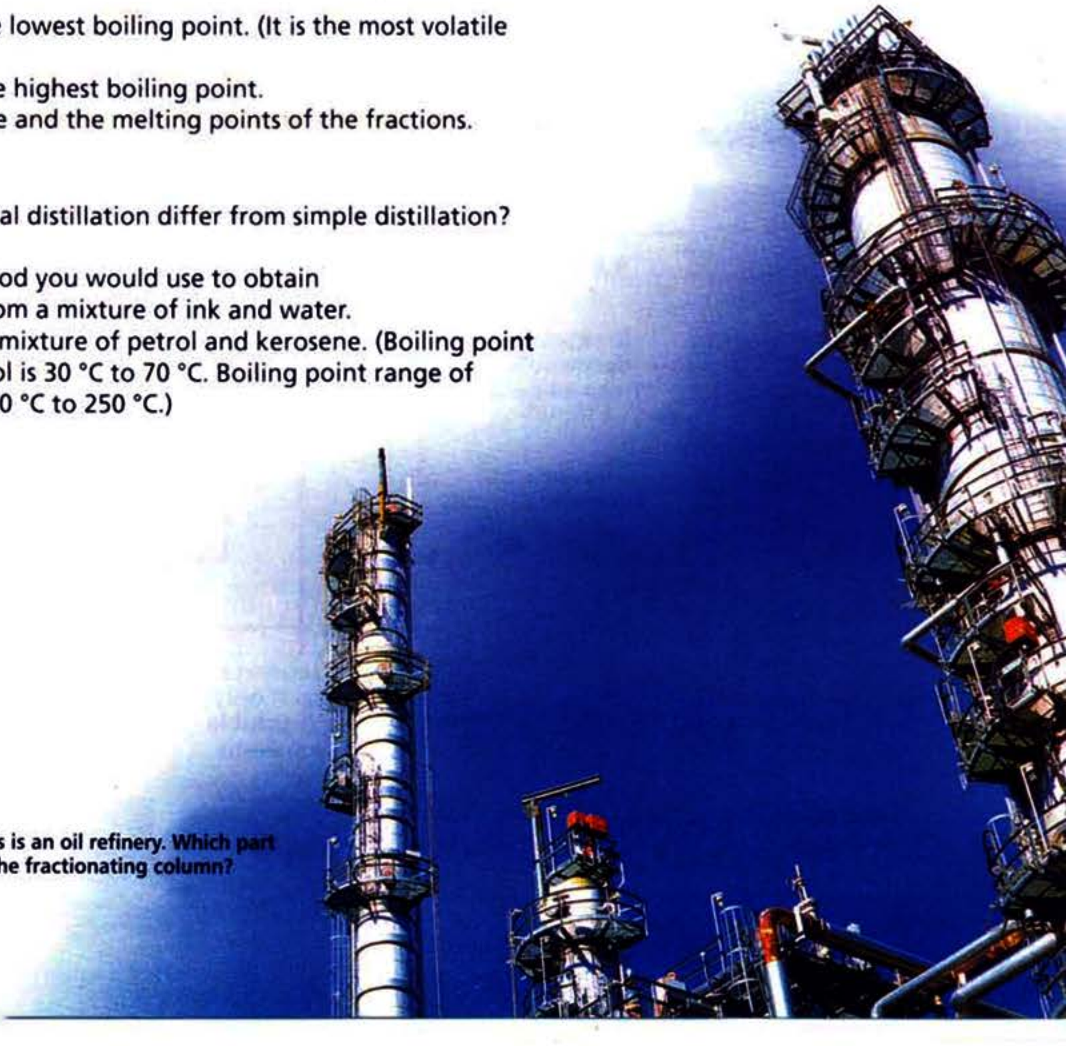
Answer

a) Fuel gas. It has the lowest boiling point. (It is the most volatile fraction.)
b) Bitumen. It has the highest boiling point.
c) Room temperature and the melting points of the fractions.

Questions

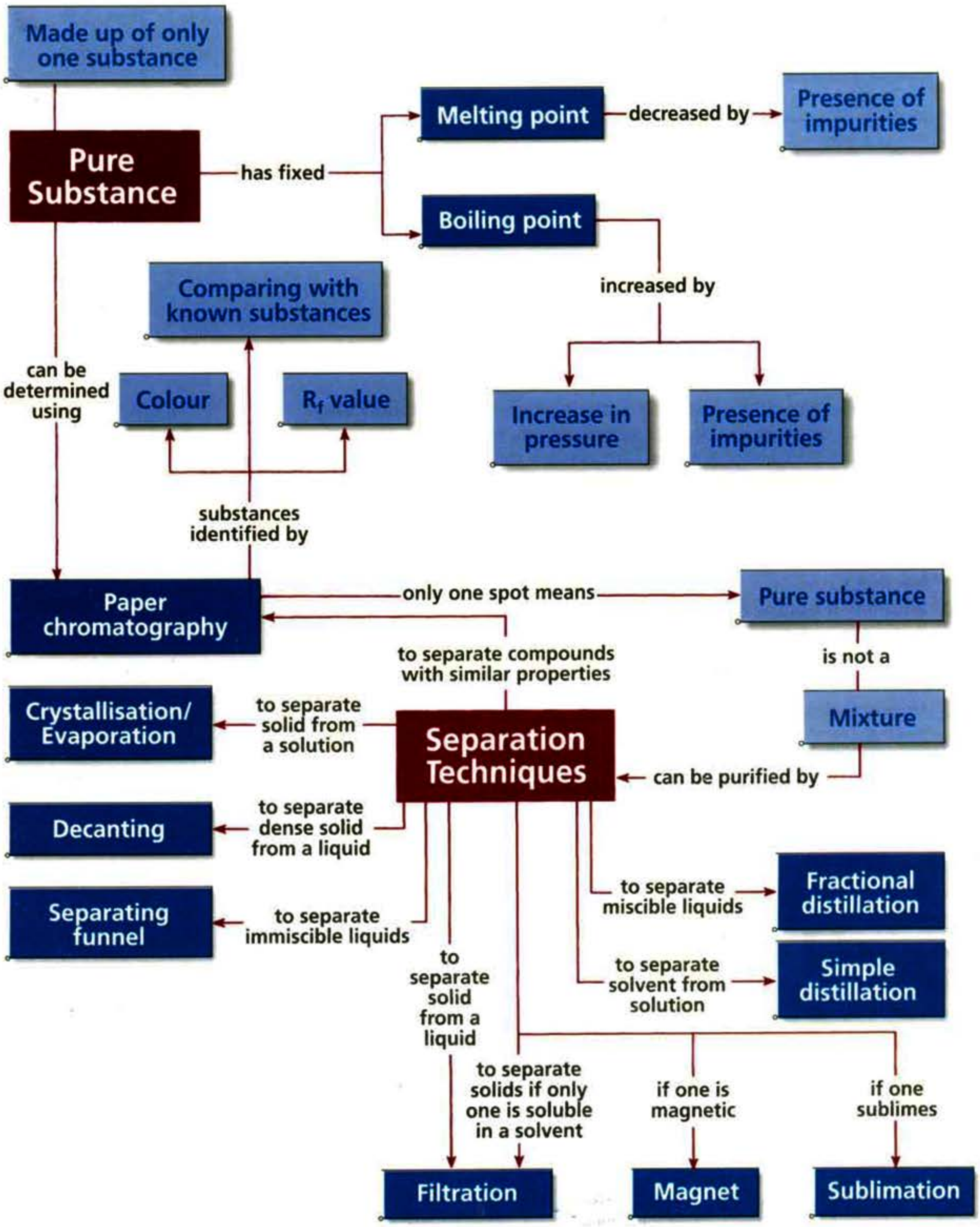
1. How does fractional distillation differ from simple distillation?

2. Describe the method you would use to obtain
   a) pure water from a mixture of ink and water.
   b) petrol from a mixture of petrol and kerosene. (Boiling point range of petrol is 30 °C to 70 °C. Boiling point range of kerosene is 160 °C to 250 °C.)



This is an oil refinery. Which part is the fractionating column?

Concept Map

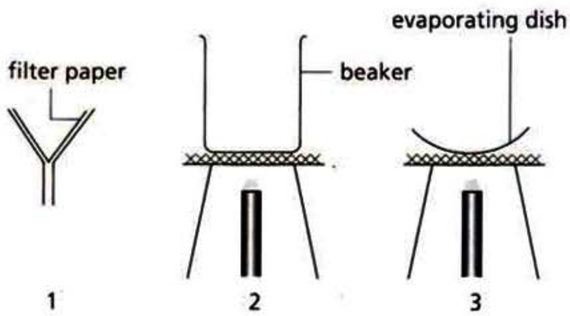

Made up of only one substance
Pure Substance
has fixed
Melting point
decreased by
Presence of impurities
Boiling point
increased by
Increase in pressure
Presence of impurities
Comparing with known substances
can be determined using
Colour
R_f value
substances identified by
Paper chromatography
only one spot means
Pure substance
is not a
Mixture
can be purified by
to separate compounds with similar properties
Separation Techniques
to separate solid from a solution
Crystallisation/ Evaporation
to separate dense solid from a liquid
Decanting
to separate immiscible liquids
Separating funnel
to separate solid from a liquid
Filtration
to separate solids if only one is soluble in a solvent
to separate miscible liquids
Fractional distillation
to separate solvent from solution
Simple distillation
if one is magnetic
Magnet
if one sublimes
Sublimation

Exercise 3

Foundation

1. Which method is the most suitable for separating a mixture of sodium chloride and iodine?
   - **A** Crystallisation
   - **B** Filtration
   - **C** Fractional distillation
   - **D** Sublimation

2. Selecting from only the apparatus shown below, how, and in what order, could separate samples of sand and salt be obtained from a mixture of seawater and sand?



   - **A** 1 then 3.
   - **B** 1 then 2.
   - **C** 3 then 1.
   - **D** 2 then 1.

3. Which process is used to test whether a solution containing a salt is saturated?
   - **A** Add a crystal of the same salt to the solution.
   - **B** Cool the solution rapidly to see if crystals will be formed.
   - **C** Heat the solution until it boils.
   - **D** Stir the solution vigorously.

4. A student was trying to make copper(II) sulphate crystals from copper(II) oxide and dilute sulphuric acid. Instead of blue crystals, his final product was a white compound. What mistake did the student make?
   - **A** Adding excess copper(II) oxide.
   - **B** Heating the filtrate to dryness.
   - **C** Using concentrated sulphuric acid instead of dilute sulphuric acid.
   - **D** Using excess dilute sulphuric acid.

5. Which process is used to separate
   a) three water-soluble dyes?
   b) two miscible liquids with boiling points of 78 °C and 100 °C?
   c) water containing an insoluble solid?
   d) water containing a dissolved solid

6. A mixture contains the following three liquids that are completely miscible:

Liquid	Boiling point (°C)
propanone	48
ethanol	78
water	100

   a) The liquids can be separated by fractional distillation. Draw a labelled diagram of the fractional distillation apparatus.
   b) State, with a reason, which liquid will distil over first.
   c) Name an industrial process that involves fractional distillation.

7. A mixture contains naphthalene, calcium fluoride and potassium chloride. The table below gives some of their properties.

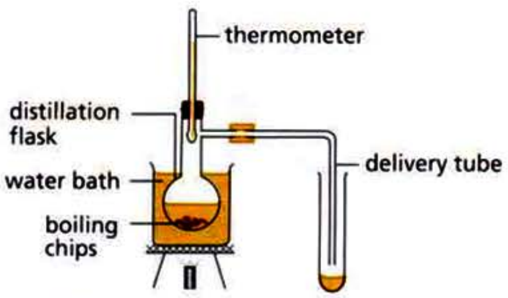
Compound	Heat	Cold water	Hot water
calcium fluoride	no effect	insoluble	insoluble
naphthalene	sublimes	insoluble	insoluble
potassium chloride	no effect	fairly soluble	very soluble

Describe how you would obtain dry and pure samples of all three compounds from the mixture.

8. Describe briefly how you would obtain each of the underlined substances.
   a) Small pieces of pure and dry iron from a mixture of iron pieces and sand.
   b) Pure and dry solid sodium chloride from a mixture of solid sodium chloride and glass.
   c) Pure water from a mixture of water and ink.

Challenge

1. Which substance, A, B, C or D, could be distilled using the apparatus below?



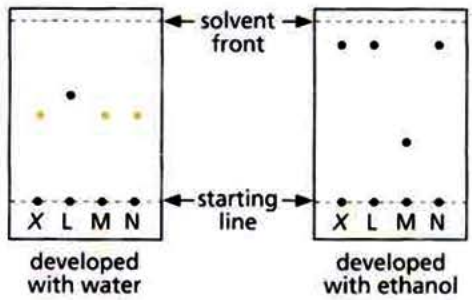
	Melting point (°C)	Boiling point (°C)
A	–138	0
B	–123	50
C	0	108
D	41	182

Problem-solving strategy:
The process is distillation. The substance must undergo two processes: boiling and condensation.

Since a water bath is used, the maximum temperature that the substance can reach is 100 °C. Which are the substances that can boil when placed in a water bath?

The boiling point is also the temperature at which a substance condenses. Which substance is able to condense at room temperature?

2. You are given an unknown substance *X* and told that it could be one of the three substances, L, M or N. You developed the two sets of chromatograms below. From the chromatograms, deduce the identity of *X*.

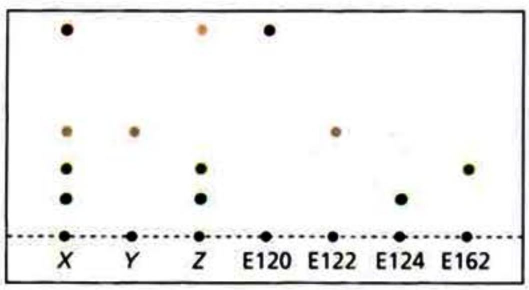


A It could be either L or M.
B It could be either M or N.
C It must be L.
D It must be N.

Problem-solving strategy:
If two substances are identical, they will have the same height on a chromatogram regardless of the solvent used.

3. Dyes that are allowed in foods are given E numbers. Four such E numbers commonly used in soups are E120, E122, E124 and E162.

   The chromatogram below was obtained from three tomato soups *X*, *Y* and *Z*. Soup *Z* was a new soup on the market. A chemist was testing it to ensure that all of the dyes used in making the soup were legal.



   a) Explain why water can be used as a solvent.
   b) Why was the pencil line, on which the spots were placed, drawn above the level of the solvent?
   c) Which soup contains all four E numbers?
   d) Are the E numbers in soup *Z* legal? Explain your answer.
   e) Kian Hok is allergic to E120, cochineal. Which soup(s) should he avoid?

4. Before air is fractionally distilled, dirt and other solid matter have to be removed. Carbon dioxide and water vapour are then removed.

   a) Suggest how dirt and other solid matter are removed.

   b) Suggest how the following are removed from the air.
      i) Carbon dioxide
      ii) Water vapour

   The remaining gases and their boiling points are given below.

Gas	Boiling point (K)
helium	4
neon	27
nitrogen	77
argon	87
oxygen	90
krypton	116
xenon	165

   Before distillation, the air is cooled to 73 K. This liquefies most of the gases. The gases are then separated by fractional distillation of liquid air.

   c) Which gases will **not** condense at 73 K?

   d) If the temperature is slowly raised from 73 K to 170 K, which gas will vaporise first?

   The gases are separated into two fractions. The first fraction is collected as the air is heated until it reaches 88 K. The second fraction contains all the other gases in the air.

   e) i) Which gases will be found in the first fraction?
      ii) Which gases will be found in the second fraction?

Chemistry Today

The Dead Sea is situated between Jordan and Israel. It is called the Dead Sea because almost nothing lives in it. However, it contains many chemicals that are useful for industry and agriculture. These chemicals are present in high concentrations in the water. In fact, the concentration of chemicals in the Dead Sea makes the water so dense that people can easily float on it!

The main chemicals extracted from the Dead Sea are magnesium chloride, magnesium bromide, potassium chloride and sodium chloride.

These chemicals can be separated by a process known as fractional crystallisation because
- they have different solubilities,
- the rate at which solubility changes with temperature is different for each chemical.

For example, sodium chloride is only slightly more soluble in hot water than in cold water. On the other hand, potassium nitrate is a lot more soluble in hot water than in cold water.

By carefully controlling the rate of evaporation and cooling, it is possible to separate the different chemicals.

CRITICAL THINKING

You are given 200 g of a mixture containing equal amounts of potassium chloride and potassium nitrate. How would you obtain pure samples of potassium chloride and potassium nitrate from this mixture?

The mass of each chemical that dissolves in 100 g of water, at three different temperatures, is given in the table below. State any assumptions that you make.

Chemical	Mass in 100 g of water at		
	20 °C	50 °C	80 °C
potassium chloride	35 g	42 g	53 g
potassium nitrate	30 g	70 g	150 g

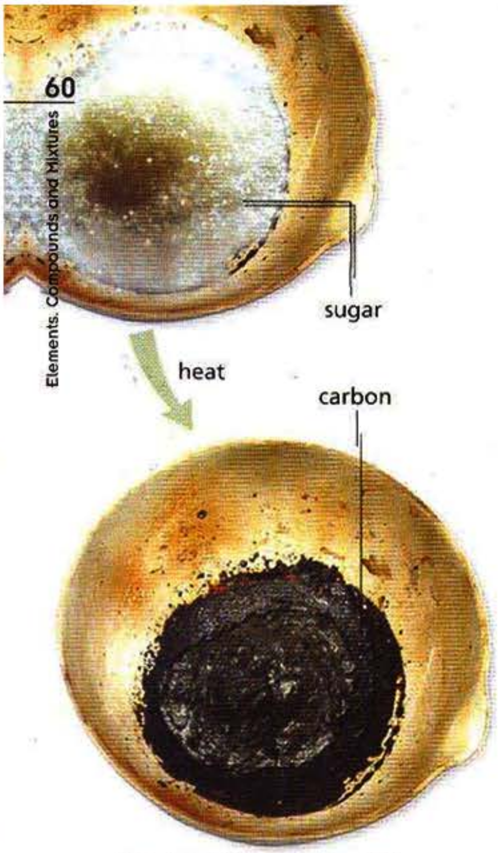
Chapter 4

Elements, Compounds and Mixtures

There are only 26 letters that make up the English alphabet. Yet, when these letters are combined in different ways, thousands of different words can be formed.

Just as letters are used to 'build' words, elements 'build' compounds. At present, 116 elements are known — 92 occur naturally, while the rest are man-made. These elements can combine to form thousands of different compounds. In chapter 1, we classified matter into three states — solid, liquid and gas. In this chapter, we shall classify matter as elements, compounds and mixtures.

Chapter Outline



Sugar can be broken down by heating. Sugar is not an element.

Chem-Aid

Quite often, you can guess the symbol from the first two letters of the element's name.
Neon Ne
Calcium Ca

What is an element?

An **element** is a *pure substance that cannot be split up into two or more simpler substances by chemical processes or by electricity*. For example, copper cannot be split into simpler substances. Therefore, copper is an element.

Sugar is not an element because it can be broken down into water and carbon. Even water is not an element. Water can be broken down by electricity to give the elements hydrogen and oxygen. Carbon, hydrogen and oxygen are elements because they cannot be further broken down into simpler substances.

Chemical Symbols of Elements

Chemists use **chemical symbols** to represent elements. Each element has a unique symbol consisting of one or two letters. For example, the symbol for oxygen is O and the symbol for iron is Fe. Table 4.1 gives the names and symbols of some common elements.

Element	Symbol	Element	Symbol
calcium	Ca	mercury	Hg
carbon	C	neon	Ne
hydrogen	H	silicon	Si
iron	Fe	sodium	Na

Table 4.1 Names and symbols of some common elements. All known elements and their symbols are recorded in a table known as the Periodic Table (see chapter 16).

Classification of Elements — Metals and Non-metals

Elements can be classified into two major groups — **metals** and **non-metals**. Iron is a metal. Oxygen is a non-metal. The different physical properties of metals and non-metals are shown in Table 4.2.

Some elements *have properties of both metals and non-metals*. Silicon is one such element. It exists as a solid and is shiny like a metal. However, it is also brittle like a non-metal. Elements like silicon, germanium, boron and arsenic are called **metalloids**. You will learn more about such elements in chapter 16.

This is a piece of charcoal. It is made entirely of one element. What element is this? Find out in chapter 7.

Metals	Non-metals
• shiny appearance (lustrous)	• dull appearance if solid (non-lustrous)
• solids at r.t.p.* (except for mercury)	• either gases, volatile liquids or solids with low melting points at r.t.p.* (except for carbon)
• malleable (can be hammered into different shapes without breaking) • sonorous (make a ringing sound when struck) • ductile (can be drawn into wires)	• brittle if solid (easily broken when hammered)
• high melting points and boiling points (except for sodium, potassium and mercury)	• low melting points and boiling points (except for carbon and silicon — see chapter 7)
• good conductors of heat	• poor conductors of heat (except carbon in the form of diamond and graphite — see chapter 7)
• good conductors of electricity in all states of matter	• poor conductors of electricity (except carbon in the form of graphite — see chapter 7)

* r.t.p. refer to the conditions of room temperature and pressure.

Table 4.2 Physical properties of metals and non-metals

Atoms and Molecules

Elements are made up of tiny particles called **atoms**. The atoms of one element are not the same as the atoms of another element.

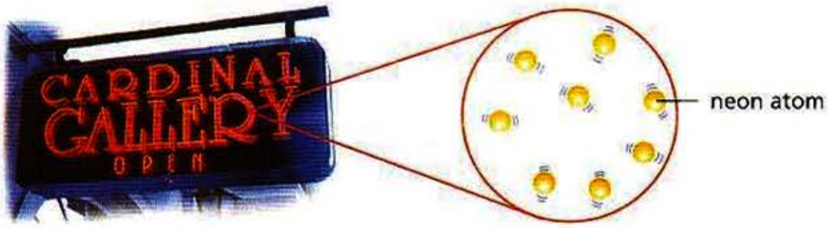


Fig. 4.1(a) The element neon is used as a gas in advertising signs. It consists entirely of neon atoms.

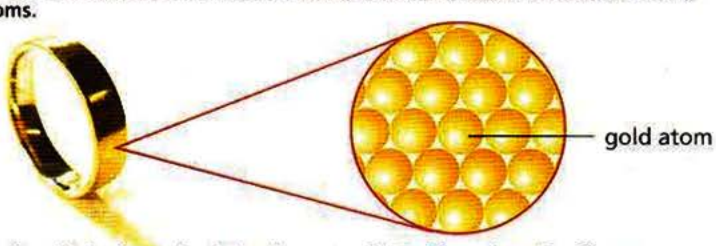
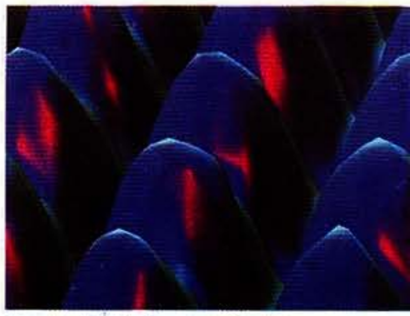


Fig. 4.1(b) A gold ring is made of the element gold. Gold consists of gold atoms, which are different from neon atoms.

Atoms are *the smallest particles of an element that have the chemical properties of that element.*

To show one atom of an element, chemists use its chemical symbol. For example, to show one atom of hydrogen, chemists write 'H'. Atoms can be split but the particles obtained no longer have the properties of the individual element. In recent years, chemists have discovered that it is possible to use electron microscopes to take pictures of atoms. In these pictures, the atoms are magnified millions of times and appear fairly spherical. You will learn more about atoms in chapter 5.



Atoms of nickel, seen using an electron microscope. We can represent each nickel atom by writing 'Ni'.

Do all elements exist as atoms?

Helium, neon, argon, krypton, xenon and radon are the only elements that exist as individual atoms. We say that they are **monatomic elements**. This means that their *atoms are not joined together chemically*. Metals also exist as atoms but large numbers of their atoms are arranged in a special way to form giant structures. Most non-metals exist as molecules. You will learn more about this in chapter 7.

What is a molecule?

A **molecule** *is a group of two or more atoms that are chemically combined (joined together)*. For example, a hydrogen molecule is made up of two hydrogen atoms chemically combined. Chemists use the symbol H_2 to represent a hydrogen molecule. We say that H_2 is the **molecular formula** of hydrogen. A molecule of ozone is formed when three atoms of oxygen are chemically combined. Thus, ozone has the molecular formula O_3.

A hydrogen molecule (H_2) An ozone molecule (O_3)

Fig. 4.2 Hydrogen and ozone molecules

Molecules *formed by the combination of two atoms* are called **diatomic molecules**. Molecules *consisting of three atoms* are called **triatomic molecules**. Molecules that are *formed by the combination of four or more atoms* are called **polyatomic molecules** (Table 4.3).

<table>
<tr><th colspan="6">Monatomic elements</th></tr>
<tr><td></td><td></td><td></td><td></td><td></td><td></td></tr>
<tr><td>Name</td><td>helium</td><td>argon</td><td>krypton</td><td>xenon</td><td>radon</td></tr>
<tr><td>Symbol</td><td>He</td><td>Ar</td><td>Kr</td><td>Xe</td><td>Rn</td></tr>
<tr><th colspan="6">Diatomic molecules</th></tr>
<tr><td></td><td></td><td></td><td></td><td></td><td></td></tr>
<tr><td>Name</td><td>nitrogen</td><td>oxygen</td><td>chlorine</td><td>bromine</td><td>iodine</td></tr>
<tr><td>Molecular formula</td><td>N_2</td><td>O_2</td><td>Cl_2</td><td>Br_2</td><td>I_2</td></tr>
<tr><th colspan="6">Polyatomic molecules</th></tr>
<tr><td></td><td colspan="2"></td><td colspan="2"></td><td></td></tr>
<tr><td>Name</td><td colspan="2">ozone</td><td colspan="2">phosphorus</td><td>sulphur</td></tr>
<tr><td>Molecular formula</td><td colspan="2">O_3</td><td colspan="2">P_4</td><td>S_8</td></tr>
</table>

Table 4.3 Atoms and molecules

Link

1. Why do the elements helium, neon, argon, krypton, xenon and radon exist as individual atoms?
2. Why do atoms of some elements combine to form molecules?

Find out in chapters 6 and 7.

Chem-Aid

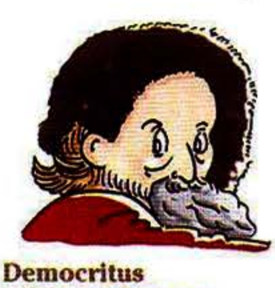
You should make sure that you understand the difference between N_2 and 2N. N_2 represents a molecule of nitrogen; 2N represents two atoms of nitrogen.

Two atoms of nitrogen (2N) One molecule of nitrogen (N_2)

Thus N_2, and not 2N, is the molecular formula of nitrogen.

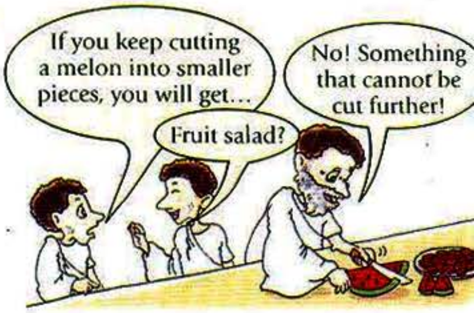
Key ideas

1. An element is a pure substance that cannot be split into simpler substances by chemical processes or by electricity.
2. Every element is represented by a unique symbol. The symbol may consist of one or two letters.
3. Elements can be classified as metals or non-metals according to their properties.
4. The atom is the smallest particle of an element that has the chemical properties of the element.
5. Metals exist as atoms in a giant structure. Most non-metals exist as molecules.
6. A molecule is made up of two or more atoms that are chemically combined.



Democritus
(460 BC – 370 BC)

Democritus was a Greek philosopher and mathematician. He is best known for his atomic theory. Democritus pondered over a key question — what would happen if one kept cutting matter into smaller and smaller pieces? What do you think?



Democritus concluded that one would eventually end up with something that was no longer divisible. He called this particle *atomos* — 'not able to be cut' or 'indivisible'. Amazingly, his theory of the existence of atoms was not proved until the beginning of the 20th century.

Test Yourself 4.1

Worked Example

Why is potassium classified as a metal?

A It is a pure substance.
B It is a good conductor of electricity.
C It exists as atoms, not molecules.
D It is very soft.

Thought Process

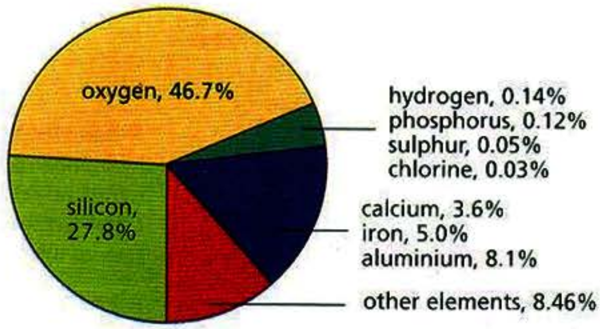
Recall the properties of metals that distinguish them from non-metals. All metals are good conductors of electricity. The only common non-metal that conducts electricity is graphite.

Answer

B

Questions

1. What are the symbols for the following pairs of elements? You may use the Periodic Table at the back of the book to find out.
   a) Magnesium and manganese.
   b) Sulphur and silicon.
   c) Zinc and tin.
2. The pie chart on the right shows the abundance of the elements by mass in planet *X*. Use the Periodic Table at the back of the book to write down the chemical symbols of these elements.



3. An element has a melting point of 120 °C and it is a non-conductor of electricity. Based on this information, predict **three** other possible physical properties of this element.
4. Give **three** reasons why pure copper is a good choice for making electrical wires.

A cake is made by baking flour, eggs, water and sugar. After baking, the end product is nothing like the starting materials. Similarly, a compound is very different from the elements that form it.

4.2 | Compounds

Combining Elements to Form a Compound

When the element magnesium burns in the element oxygen, a brilliant white flame is seen and a new substance, magnesium oxide, is obtained. The new substance, magnesium oxide, is called a **compound**. It has very different properties from the elements that form it.

magnesium + oxygen ⟶ magnesium oxide

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

A compound is *a pure substance that contains two or more elements chemically combined.* Table 4.4 shows a list of common compounds and the elements they contain.

Compound	Elements present
common salt (sodium chloride)	sodium, chlorine
carbon dioxide	carbon, oxygen
marble (calcium carbonate)	calcium, carbon, oxygen
copper(II) sulphate	copper, sulphur, oxygen
hydrogen chloride	chlorine, hydrogen

Table 4.4 Common compounds and the elements they contain. Note that compounds have a chemical name indicating the elements present in them.

Chem-Aid

1. A positively charged ion is formed when an atom loses electrons.
2. A negatively charged ion is formed when an atom gains electrons.
3. Some ions are made up of groups of atoms. They are called polyatomic ions. Polyatomic ions can be positively or negatively charged.
4. Ions with opposite charges can combine to form compounds.

Naming Compounds

Compounds are named according to some general rules. Use the rules below to help you identify or name compounds.

General rule for naming compounds	Examples
A compound made up of two elements has a name that ends in **-ide**.	• **Sodium chlor*ide*** — made up of the elements sodium and chlorine • **Zinc ox*ide*** — made up of the elements zinc and oxygen • **Carbon diox*ide*** — made up of the elements carbon and oxygen
A compound that contains hydroxide ions, OH^- (a negatively charged ion made up of oxygen and hydrogen) is named a **hydroxide.**	• **Potassium *hydroxide*** — contains potassium ions and hydroxide ions
A compound that contains a negatively charged polyatomic ion containing oxygen usually has a name ending in **-ate.**	• **Copper(II) sulph*ate*** — contains four oxygen atoms in one sulphate ion • **Sodium nitr*ate*** — contains three oxygen atoms in one nitrate ion

Table 4.5 General rules for identifying or naming compounds

However, there are compounds with names ending in -ide that contain more than two elements. To name compounds that contain more than two elements, there are other rules to note.

Not all compounds with names ending in -ate contain negatively charged polyatomic ions containing oxygen. For these compounds, other rules are used to name them.

Fixed Composition of Compounds

A compound is made up of different elements chemically combined in a fixed ratio. For example, water (H_2O) is a compound made only by joining together two atoms of hydrogen to one atom of oxygen. That is, the ratio of hydrogen atoms to oxygen atoms in water is always 2 : 1.

Words have a fixed spelling and changing any one of the letters would give a different word altogether. Similarly, compounds have a fixed composition. Changing any one element in a compound would give a different compound altogether.

What is the smallest particle of a compound?

You have learnt that the smallest particle of an element that can exist independently is the atom. A compound is made up of at least two elements. Thus, the smallest particle of a compound must contain at least two atoms. The smallest particle of a compound that can exist independently is therefore the **molecule**. Sometimes, depending on the type of compound, the smallest particle is called the **formula unit**. You will learn the difference between a molecule and a formula unit in chapter 6.

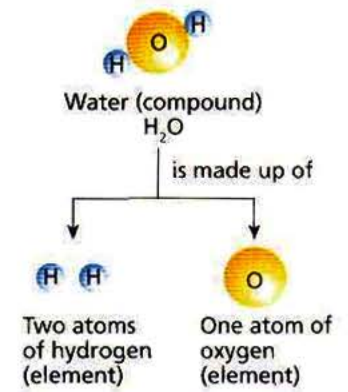


Fig. 4.3 The ratio of hydrogen to oxygen in water is always 2 : 1.

Chemical Formula of a Compound

A compound can be represented by a **chemical formula**. The chemical formula of a compound is written by putting together the chemical symbols of the elements that make up the compound.

The chemical formula states
- the types of atoms (i.e. elements) present in the compound,
- the ratio of the different atoms present in the compound.

For example, look at the chemical formula of water in Fig. 4.4. It states that
- hydrogen and oxygen atoms are present in the compound,
- the ratio of hydrogen to oxygen is 2 : 1.

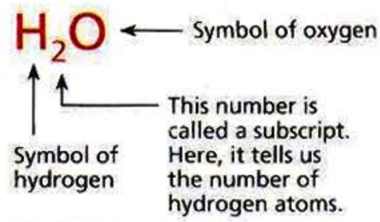


Fig. 4.4 Chemical formula of water

There are some general rules to follow when we write the chemical formula of a compound.

General rule	Examples
For many compounds that contain both metallic and non-metallic elements, the symbol of the metallic element is written first.	• calcium oxide (CaO) • sodium chloride (NaCl) • magnesium carbonate ($MgCO_3$)
The number of atoms is written as a subscript, to the right of the atom's symbol.	• water (H_2O, not H2O or $_2HO$) • magnesium carbonate ($MgCO_3$, not MgCO3 or MgC_3O)
It is not necessary to write the subscript '1'.	• water (H_2O, not H_2O_1) • calcium oxide (CaO, not Ca_1O_1)
The oxygen atom is usually written at the end of the formula.	• water (H_2O, not OH_2) • carbon dioxide (CO_2, not O_2C) • nitric acid (HNO_3, not O_3NH)

Table 4.6 General rules for writing the chemical formula of a compound

This rule does not apply for a group of compounds known as organic compounds. Find out more about organic compounds in chapter 21.

Compound	Chemical formula	Ratio of atoms
hydrogen chloride	HCl	H : Cl = 1 : 1
carbon dioxide	CO_2	C : O = 1 : 2
sulphuric acid	H_2SO_4	H : S : O = 2 : 1 : 4
ethanol	C_2H_5OH	C : H : O = 2 : 6 : 1

Table 4.7 The chemical formulae of some common compounds

The chemical formulae of some common substances are given in Table 4.7. Notice the fixed ratio of atoms.

Some chemical formulae contain brackets. For example, the formula of lead(II) nitrate is $Pb(NO_3)_2$. How then is the ratio of lead, nitrogen and oxygen atoms in lead(II) nitrate calculated? Study Fig. 4.5 to find out.

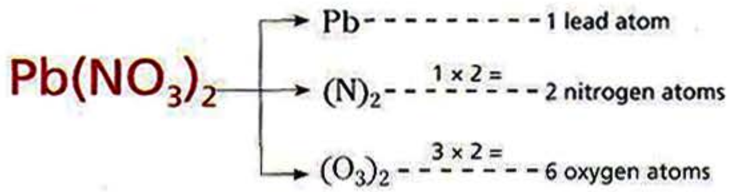


Fig. 4.5 Finding the ratio of atoms in lead(II) nitrate

Thus the ratio of atoms in lead(II) nitrate is lead : nitrogen : oxygen = 1 : 2 : 6.

Decomposition of Compounds

Heat can be used to form compounds. *Heat can also be used to break down compounds into elements or simpler compounds.* Such a chemical reaction is called **thermal decomposition.**

For example, when mercury(II) oxide is heated strongly, it decomposes to give the elements mercury and oxygen.

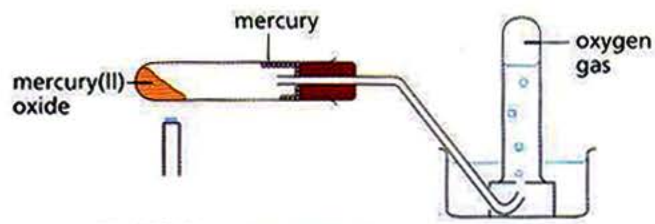


Fig. 4.6 Thermal decomposition of mercury(II) oxide

Besides using heat, compounds can also be broken down into simpler substances by using electricity.

Quick Check

What is the ratio of hydrogen gas to oxygen gas if water is decomposed?

Link

Why are some compounds extremely difficult to break down? Find out in chapter 6.

Key ideas

1. A compound is a pure substance consisting of two or more elements that are chemically combined.
2. The different elements present in a compound are combined in a fixed ratio.
3. Every compound has a unique chemical formula.
4. The chemical formula of a compound gives information on
   - the types of atoms (elements) present in the compound,
   - the ratio of different atoms present in the compound.
5. A compound can only be decomposed by chemical reactions or electricity.

Test Yourself 4.2

Worked Example

Write the full name of the following compounds. You may refer to the Periodic Table at the back of the book to complete your answers.

a) MgO   b) $CaCO_3$

Thought Process

a) MgO is made up of two elements and thus its full name ends in –ide. (Mg — magnesium, O — oxygen)

b) $CaCO_3$ is made up of three elements, including oxygen. Thus, its full name ends in –ate. (Ca — calcium, C — carbon, O — oxygen)

Answer

a) Magnesium oxide   b) Calcium carbonate

Questions

1. The chemical formulae of several substances are listed below:

   AlF_3, Co, CoO, CO, $NaBrO_3$, Na_2S, V_2O_5, $CaSO_4$, Cu

   a) Classify the substances as either (i) elements or (ii) compounds.
   b) Name the elements present in each of the compounds listed.
   c) Which substances have names ending in (i) –ide, (ii) –ate?

2. How many hydrogen atoms are present in one molecule/formula unit of each of the following compounds?

   a) Ammonia, NH_3
   b) Calcium chloride, $CaCl_2$
   c) Nitric acid, HNO_3
   d) Urea, $CO(NH_2)_2$
   e) Hydrated copper(II) sulphate, $CuSO_4.5H_2O$
   f) Ammonium sulphate, $(NH_4)_2SO_4$

3. The chemical formula of potassium dichromate(VI) is $K_2Cr_2O_7$. State the ratio of potassium, chromium and oxygen atoms in the compound.

4. Glass is mainly made up of sodium silicate. Sodium silicate contains sodium, oxygen and silicon. Its chemical formula is represented as X_2YZ_3, where *X*, *Y* and *Z* are unknown. Replace *X*, *Y* and *Z* with the correct chemical symbols they represent.

Chem-Aid

Note the difference between the chemical symbols — Co for cobalt and CO for carbon monoxide; Mn is manganese and Mg is magnesium.

Note also how *fluorine* is spelled.

4.3 | Mixtures

What are mixtures?

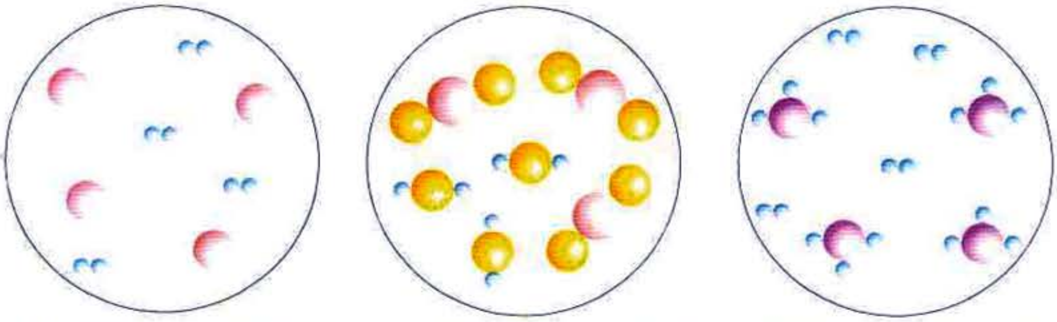
Mixtures are formed when *two substances are added together without chemical bonds being formed.* Examples of mixtures include muddy water and air. Air is made up of gases such as nitrogen and oxygen mixed together.

Link

Learn more about chemical bonding in chapters 6 and 7.

Mixtures can be made up of elements and compounds. The components of a mixture are not fixed. They can be present in any ratio. Fig. 4.7 shows the arrangement of particles in

a) a mixture of two elements,
b) a mixture of two compounds,
c) a mixture of one element and one compound.



(a) A mixture of two elements, e.g. neon (Ne) and hydrogen (H_2)

(b) A mixture of two compounds, e.g. water vapour (H_2O) and carbon dioxide (CO_2)

(c) A mixture of one element and one compound, e.g. hydrogen (H_2) and ammonia (NH_3)

Fig. 4.7 Arrangement of particles in different mixtures

If we add sulphur powder to iron filings, a mixture of iron and sulphur is formed. However, when the same mixture of iron filings and sulphur powder is heated, the mixture glows and a black solid, iron(II) sulphide, is obtained. Iron(II) sulphide is a compound of two elements, iron and sulphur. It has entirely different properties from the unheated mixture.

Quick Check

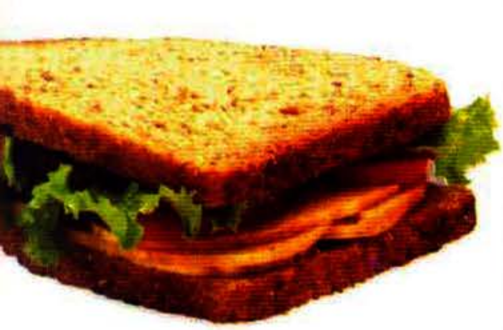
Air is a mixture of gases. Name **four** gases present in air. Give a property of air that shows that it is a mixture.

What are the differences between a mixture and a compound?

A mixture is different from a compound in a number of ways. Table 4.8 summarises the general differences between a mixture and a compound.

	Mixture	Compound
Separation	The components of a mixture can be separated by physical methods, e.g. filtration, distillation or chromatography.	The elements in a compound can only be separated by chemical reactions or by using electricity.
Properties	The chemical properties of a mixture are the same as those of its components.	The physical and chemical properties of a compound are different from those of the elements in the compound.
Energy change	No chemical reaction takes place when a mixture is formed — usually there is little or no energy change.	A chemical reaction takes place when a compound is formed — usually there is an energy change, e.g. the reactants get hot.
Composition	The components of a mixture can be mixed in any proportion.	The elements in a compound are always combined in a fixed proportion (by mass).

Table 4.8 Differences between a mixture and a compound



A mixture can be separated by physical means, unlike a compound. For example, the ingredients that make a sandwich can be separated rather easily — the bread, vegetables and meat are not chemically combined, unlike the ingredients of a cake.

Alloys

What is an alloy?

An **alloy** is an example of a mixture. An alloy is *a mixture of metals with other elements* (usually metals but sometimes non-metals such as carbon). Alloys are widely used and in great demand because they tend to be stronger than pure metals. Table 4.9 shows some common alloys and their compositions.

Alloy	Composition (percentage of element by mass)
steel*	iron (≈99%), carbon (≈1%)
stainless steel*	iron (73%), chromium (18%), nickel (8%), carbon (1%)
brass	copper (66 – 70%), zinc (30 – 34%)
bronze	copper (90%), tin (10%)
duralumin	aluminium (95%), copper (4%), magnesium (1%)
solder	lead (67%), tin (33%)

* The percentage of iron and carbon varies for different types of steel.

Table 4.9 The composition of some alloys

Alloys have chemical properties similar to those of the elements they contain but they have different physical properties. Solder is an alloy of lead and tin. It melts at around 183 °C, lead melts at 327 °C and tin at 232 °C. You will learn more about alloys in chapter 14.

This statue is made of bronze, an alloy of copper and tin.

Key ideas

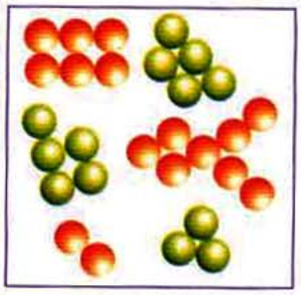
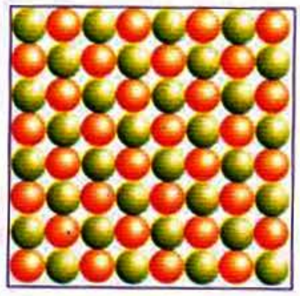
1. A mixture contains two or more substances that can be separated by physical means.
2. An alloy is a mixture of a metal with other elements (metals or non-metals).

Test Yourself 4.3

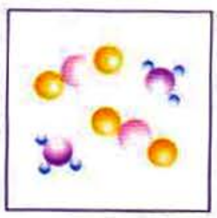
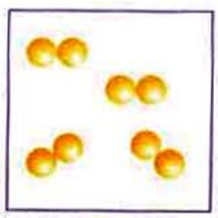
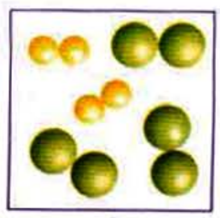
Worked Example

The elements iron and sulphur combine to form the compound iron sulphide (FeS). List the differences between iron sulphide and a mixture of iron and sulphur.

Answer

	Mixture of iron and sulphur	Compound iron sulphide
Separation	can be separated by physical means, e.g. using a magnet	can only be separated by chemical means or by using electricity
Properties	yellow specks (sulphur) mixed with dark grey powder (iron filings)	black solid
Energy change	no heat required for the mixture to form	heat required for the compound to form
Composition	ratio of iron to sulphur can vary 	ratio of iron to sulphur is always the same 

Questions

1.   

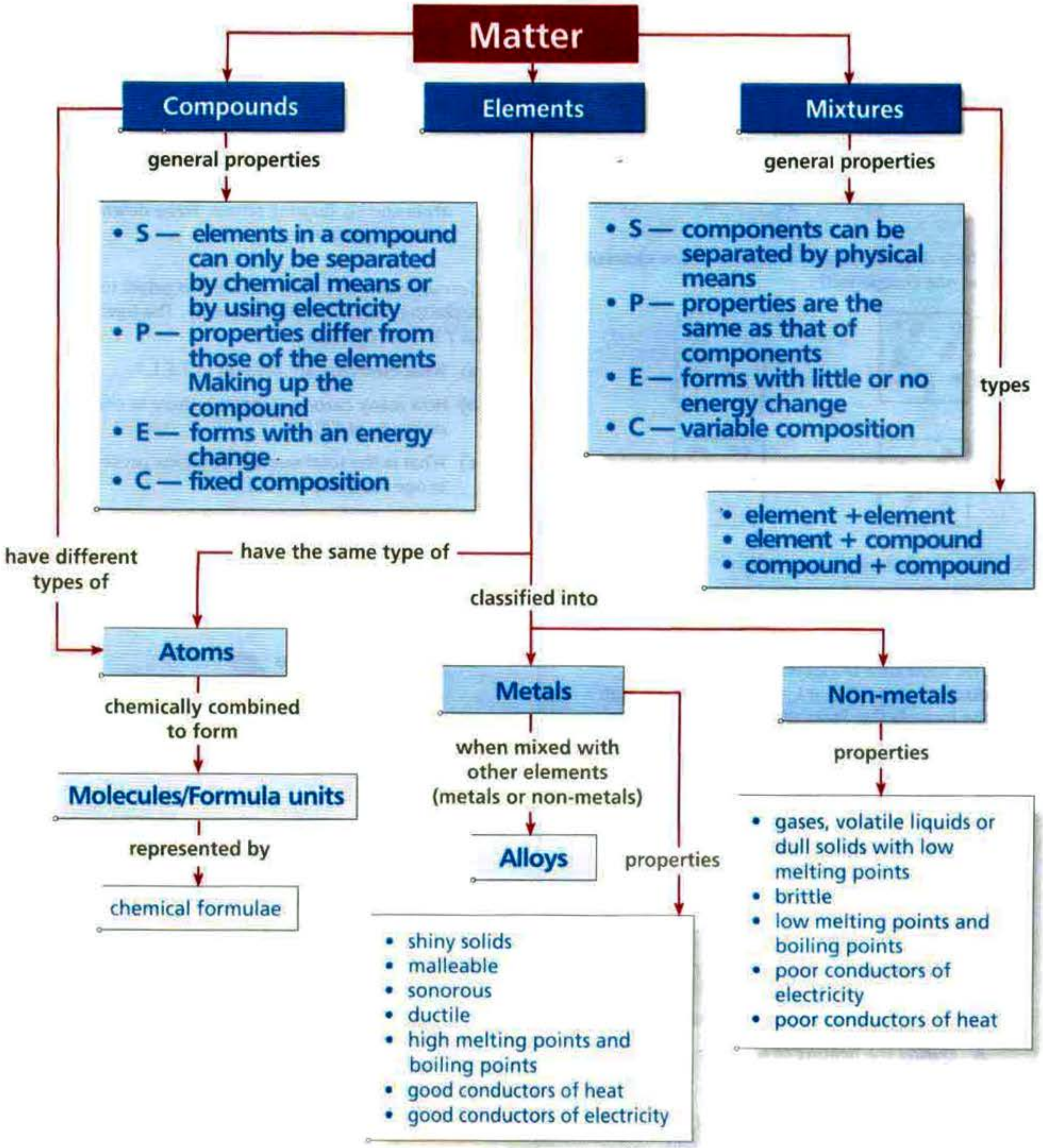
A B C

Which one of the diagrams A, B or C, represents
a) oxygen (O_2)?
b) a mixture of oxygen (O_2) and chlorine (Cl_2)?
c) a mixture of ammonia (NH_3) and carbon dioxide (CO_2)?
Explain your answers.

2. Classify the substances below as elements, compounds or mixtures. Present your answer in the form of a graphic organiser.

bromine	distilled water	duralumin	gold dust
honey	ink	magnesium chloride	

Concept Map


Matter
Compounds
Elements
Mixtures
general properties
• S — elements in a compound can only be separated by chemical means or by using electricity
• P — properties differ from those of the elements Making up the compound
• E — forms with an energy change
• C — fixed composition
general properties
• S — components can be separated by physical means
• P — properties are the same as that of components
• E — forms with little or no energy change
• C — variable composition
types
• element +element
• element + compound
• compound + compound
have different types of
have the same type of
classified into
Atoms
chemically combined to form
Molecules/Formula units
represented by
chemical formulae
Metals
when mixed with other elements (metals or non-metals)
Alloys
properties
• shiny solids
• malleable
• sonorous
• ductile
• high melting points and boiling points
• good conductors of heat
• good conductors of electricity
Non-metals
properties
• gases, volatile liquids or dull solids with low melting points
• brittle
• low melting points and boiling points
• poor conductors of electricity
• poor conductors of heat

Exercise 4

Foundation

1. Which list of symbols represents the elements fluorine, copper, magnesium and potassium respectively?

A	FL	Co	Mn	P
B	F*l*	Cu	Mg	K
C	Fe	Co	Mn	P
D	F	Cu	Mg	K

2. Which diagram shows a mixture of one element and one compound?

A

B

C

D


3. Which substance is an element?

A Bronze
B Silver
C Silver bromide
D Water

4. The substance buckminsterfullerene was discovered in 1985. It is a perfect sphere with the formula C_{60}. From this information, what can you deduce about buckminsterfullerene?

A It contains only one element.
B It is a compound of 60 elements.
C It is a mixture of 60 atoms.
D It is a mixture of 60 elements.

5. When mercury oxide is heated, it forms mercury and oxygen.

mercury oxide(s) $\longrightarrow$ mercury(*l*) + oxygen(g)

What can be deduced from this information?

A During the heating of mercury oxide, all three states of matter exist.
B Mercury oxide is a mixture.
C Mercury oxide is a silver solid.
D The composition of mercury oxide varies during heating.

6. a) Which elements are contained in each of the following compounds?
   i) Nickel(II) sulphate
   ii) Copper(II) hydroxide
   iii) Sulphur dioxide

   b) Cryolite is a compound of aluminium, sodium and fluorine. One formula unit of cryolite contains three sodium atoms, one aluminium atom and six fluorine atoms. Write down the chemical formula of cryolite.

7. Tetraethyl lead (T.E.L.) is added to petrol to make the car engine run smoothly. The formula of T.E.L. is $(C_2H_5)_4Pb$.

   a) What elements are present in T.E.L.?

   b) How many carbon atoms are there in one molecule of T.E.L.?

   c) What is the total number of atoms present in one molecule of T.E.L.?

8. The diagrams below represent molecules of ammonia (NH_3), hydrogen chloride (HCl), methane (CH_4) and water (H_2O).

   Identify A, B, C and D.

A

B

C

D


9. a) Compare the formulae of the following substances and then classify them into **three** different groups e.g. based on the elements in them. Briefly explain the reason(s) for your groupings.

   C_2H_6, $C_2H_4O_2$, CH_3NH_2, $C_3H_6O_2$, C_6H_6, NH_3

   b) Give reasons why air is considered to be a mixture.

Chemistry Today

In the past, some elements were named according to the substances they produced. These elements have names that end with 'gen' (abbreviated from 'generator'). Thus hydrogen means 'water generator' and nitrogen means 'nitric acid generator'.

Today, the names of new elements have to be approved by the International Union of Pure and Applied Chemistry (IUPAC). Names are not given to elements until the discovery of the new element has been confirmed by another laboratory. The table below gives the names of elements most recently approved, the year they were discovered, and the reason for giving the element that name. Notice how most of the elements were named after the person who discovered it. Also included in the table are elements that have not yet been officially named.

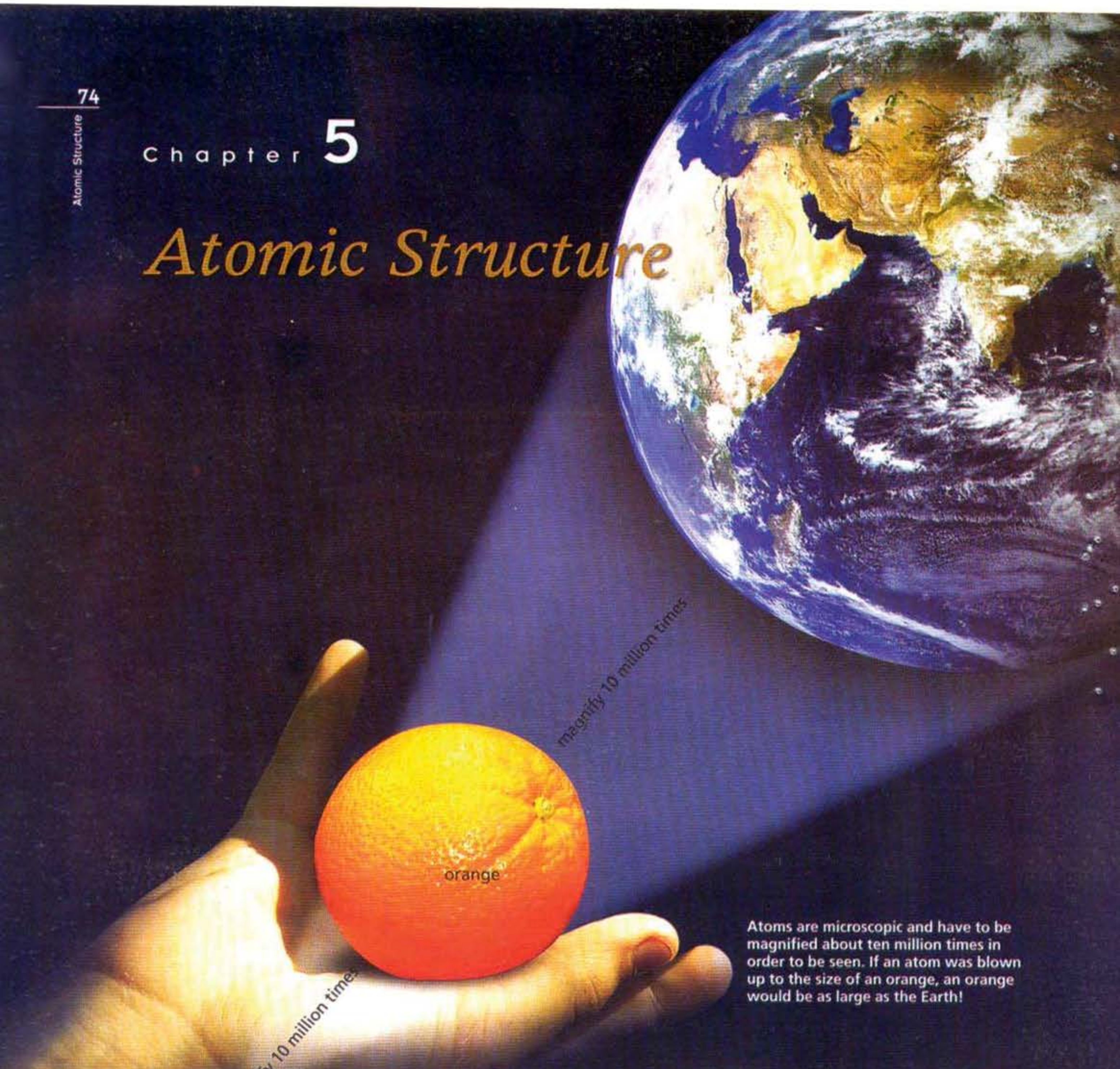
Element	Name	Symbol	Year of discovery	Named after
104	Rutherfordium	Rf	1969	Ernest Rutherford, New Zealander nuclear physicist and chemist
106	Seaborgium	Sg	1974	Glenn Seaborg, American nuclear physicist
107	Bohrium	Bh	1976	Niels Bohr, Danish physicist
109	Meitnerium	Mt	1982	Lise Meitner, Austrian physicist
110	Darmstadtium	Ds	1994	Darmstadt, Germany, place of discovery
111	Roentgenium	Rg	1994	Wilhelm Roentgen, German scientist, discoverer of X-rays
112	Ununbium	Uub	1996	not yet officially named by IUPAC
113	Ununtrium	Uut	2004	

CRITICAL THINKING

Element 118 has not been discovered yet. It is predicted that this element will be a noble gas like helium. Suggest some physical and chemical properties of element 118.

Chapter 5

Atomic Structure



Atoms are microscopic and have to be magnified about ten million times in order to be seen. If an atom was blown up to the size of an orange, an orange would be as large as the Earth!

Chapter Outline

All matter is made up of atoms. Up until the mid-1800s, most scientists thought that atoms were solid through and through. However, over the last 100 years, it has been discovered that atoms contain even smaller particles inside!

5.1 | Inside Atoms

What are atoms made up of?

Atoms are made up of three different particles — **protons**, **neutrons** and **electrons**. These particles are even smaller than the atoms themselves. We call them **sub-atomic particles**.

Look at Fig. 5.1. It shows the simplified structure of an atom. Protons and neutrons are tightly packed together in the centre of an atom, forming the **nucleus** of the atom. *Collectively, protons and neutrons are known as* **nucleons**. Electrons move rapidly around the nucleus.

Try it Out

Atoms are so small that they can only be seen using powerful microscopes that were developed in the 1980s. So how did scientists first discover what was inside the atom as early as the 1900s? Use the internet to find out more.

nucleus

Electron
Each electron
- carries one negative electric charge (–1),
- has a relative mass of $\frac{1}{1840}$ (negligible),
- is represented by the symbol, e.

Contains protons
Each proton
- carries one positive electric charge (+1),
- has a relative mass of 1,
- is represented by the symbol, p.

Contains neutrons
Each neutron
- carries no electric charge,
- has a relative mass of 1,
- is represented by the symbol, n.

Fig. 5.1 Structure of an atom

As Fig. 5.1 shows, the different sub-atomic particles have different electric charges and relative masses.

Chem-Aid

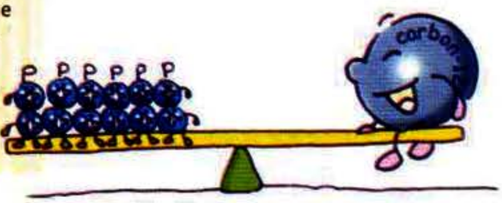
1. The particles in an atom can be remembered by **PEN** —
   Protons
   Electrons
   Neutrons
2. The relative mass of each type of sub-atomic particle is obtained by comparing its mass to the mass of a carbon atom (called carbon-12).
3. The relative mass of an electron is negligible or $\frac{1}{1840}$. It is not zero.

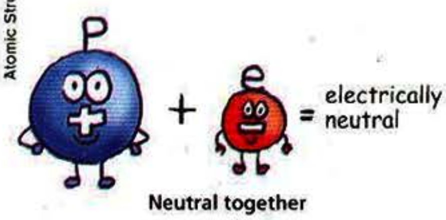
TidBit

All the particles in an atom are very light. They are so light that it is not convenient to express their masses in grams or kilograms. It is easier to measure their masses against a standard unit, called the atomic mass unit (amu).

1 amu = 1.67×10^{-27} kg
= 0.000 000 000 000 000 000 000 000 001 67 kg

1 amu is one-twelfth the mass of a carbon atom (called carbon-12). The carbon-12 atom has six protons, six neutrons and six electrons. The mass of this carbon atom was determined accurately by experiment. Protons and neutrons have the same mass, about 1 amu each. We say that a proton (or a neutron) has a relative mass of 1. An electron has an actual mass of 9.1×10^{-31} kg, which is $\frac{1}{1840}$ that of a proton's mass. When we compare the mass of electrons to that of protons and neutrons in an atom, the mass of electrons is insignificant. Thus, the relative mass of an electron is often considered negligible.





Neutral together

Quick Check

The total number of protons and electrons in an atom of element *X* is 30. What is the atomic number of element *X*?

5.2 | The Proton Number and Nucleon Number

All atoms are electrically *neutral*. An atom contains an *equal number of positively charged protons and negatively charged electrons*. The positive and the negative electric charges cancel out exactly.

What is the proton number?

The number of protons in an atom is called the **proton number** or the **atomic number**. The proton number is represented by the symbol **Z**. Since atoms are neutral, the proton number can also tell us the number of electrons in the atom. For example, the proton number of nitrogen is 7. This implies that the nitrogen atom has seven protons and seven electrons.

Different atoms have different proton numbers

Atoms of different elements have different numbers of protons. Each element has a unique proton number. This means that no other element has atoms with this proton number. Carbon, for example, has a proton number of 6. All atoms with six protons are therefore carbon atoms. Oxygen has a proton number of 8. Thus, all atoms with eight protons are oxygen atoms.

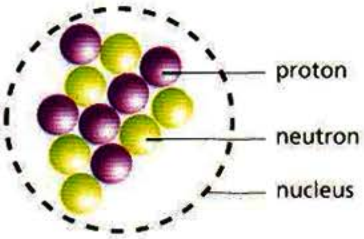


Fig. 5.2(a) This atom has six protons in its nucleus. Therefore it is a carbon atom.

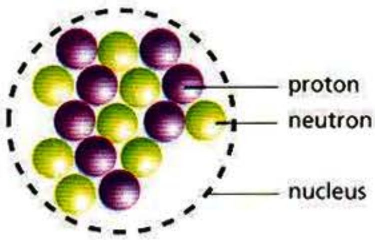
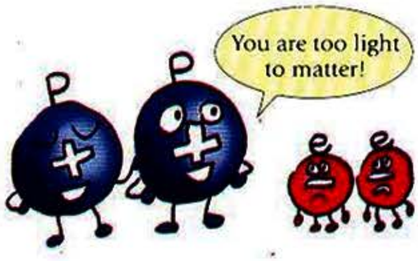


Fig. 5.2(b) This atom has eight protons in its nucleus. Therefore it is an oxygen atom

What is the nucleon number?

The total number of protons and neutrons in an atom is called the **nucleon number**. It is represented by the letter **A**. The nucleon number is also called the **mass number**. This is because the mass of an atom depends on the number of protons and neutrons in the atom's nucleus. The mass of electrons in the atom is said to be negligible.

Nucleon number (A) = number of protons + number of neutrons



The nucleon and proton numbers can be included when representing an element in symbols. This is shown below for the element sodium.

nucleon number (mass number) in superscript, left of symbol — $^{23}_{11}\text{Na}$ — chemical (atomic) symbol

proton number (atomic number) in subscript, left of symbol

Fig. 5.3 Representing an element with nucleon and proton numbers

For convenience, sometimes the element is represented using only the nucleon number e.g. sodium-23 or ^{23}Na.

Key ideas

1. An atom is made up of three types of sub-atomic particles — protons, neutrons and electrons.
2. Sub-atomic particles differ in mass and electric charge.

Sub-atomic particle	Symbol	Relative mass	Relative charge
Proton	p	1	+1
Neutron	n	1	0
Electron	e	$\frac{1}{1840}$	−1

3. Proton number, Z = number of protons
   Nucleon number, A = total number of protons and neutrons
   Number of neutrons = A − Z
4. An atom is electrically neutral. This means that in an atom, the number of protons is equal to the number of electrons.

Test Yourself 5.1

Worked Example

Deduce the number of protons, neutrons and electrons in an atom of uranium-235 ($^{235}_{92}$U).

Answer

Number of protons = 92
Number of electrons = 92
Number of neutrons (A − Z) = 235 − 92 = 143

Questions

1. Compare a neutron and an electron. How are they different?
2. Element *X* has a proton number of 7. It also has seven neutrons.
   a) Deduce the number of electrons and the nucleon number of *X*.
   b) Represent *X* by writing the chemical symbol, including the proton and nucleon numbers.

John Dalton (1766 – 1844)

Did you know that this scientist, also called the 'father of modern atomic theory', could not see the colours red, orange and green? Nevertheless, John Dalton did not let his colour-blindness get in the way of his scientific work. Many of his ideas led to the foundation of modern chemistry. Dalton was the first person to use the word 'atom' after Democritus. His ideas on the nature of atoms helped answer these questions:

- Do atoms of an element have the same mass?
- Do atoms of different elements have different masses?
- Can atoms be created or destroyed?

Chlorine is used to kill bacteria in swimming pools.

5.3 | Isotopes

Look at the models of hydrogen atoms shown in Fig. 5.4. What is the same or different about them?

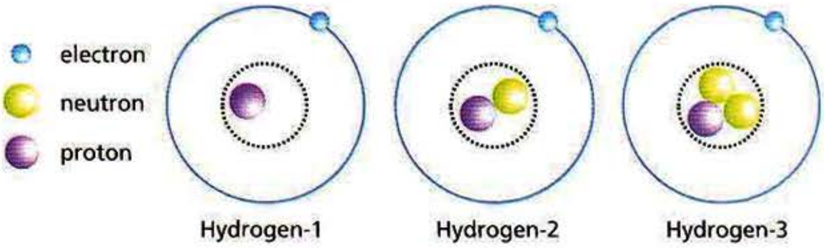


Fig. 5.4 Hydrogen atoms

They are similar except that hydrogen-1 has no neutrons, hydrogen-2 has one neutron and hydrogen-3 has two neutrons. These hydrogen atoms are known as **isotopes**.

What are isotopes?

Isotopes are *atoms of the same element with the same number of protons but different number of neutrons*. Isotopes of the same element have the same number of electrons. Thus, isotopes have the same proton number but different nucleon numbers.

Fig. 5.5 Isotopes of chlorine

Most elements that commonly occur are made up of isotopes. For example, chlorine consists of two isotopes. A sample of chlorine gas consists of 75% chlorine-35 and 25% chlorine-37. The simplified atomic structures of the isotopes are shown in Fig. 5.5. A few elements, however, do not have isotopes. For example, all atoms of fluorine contain ten neutrons and nine protons.

Isotopes have the same chemical properties but slightly different physical properties. The chemical properties of isotopes are similar because chemical reactions involve only the electrons and not the protons and neutrons. The physical properties differ because the relative masses of the isotopes differ. For example, hydrogen-2 has a slightly higher boiling point and density than hydrogen-1.

Uses of Isotopes

Isotopes that emit high-energy radiation are called **radioisotopes**. They are classified as radioactive substances. Radiation emitted by radioisotopes is dangerous because it can damage living cells and cause cancer. However, radioisotopes can have important applications and can be safely used if they are handled properly.

A heart pacemaker is implanted in patients with heart problems to regulate their heartbeats. Pacemakers are powered by lithium batteries nowadays. In the past, they were powered by the isotope plutonium-238.

Chem-Aid

Physical properties include mass, boiling and melting points, density, etc.
Chemical properties determine how an atom behaves during a chemical reaction.

5.4 | Arranging Electrons in Atoms

Between the years 1912 and 1913, scientist Lord Rutherford and two of his assistants, Geiger and Marsden, conducted experiments to find out more about the structure of atoms. What they did led to new knowledge about the arrangement of electrons in atoms.

Alpha particles are positively charged helium particles ($^{4}_{2}He^{2+}$). The scientists fired alpha particles at a very thin sheet of gold foil. Most of the alpha particles passed straight through the foil. A few were deflected and about 1 in 10 000 alpha particles bounced back from the foil. To Rutherford, the fact that a few particles were bounced right back was incredible. It was as if they had fired a bullet at a sheet of tissue paper and it came back to hit them!

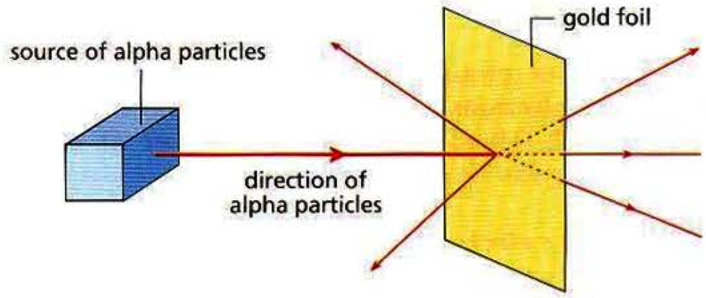
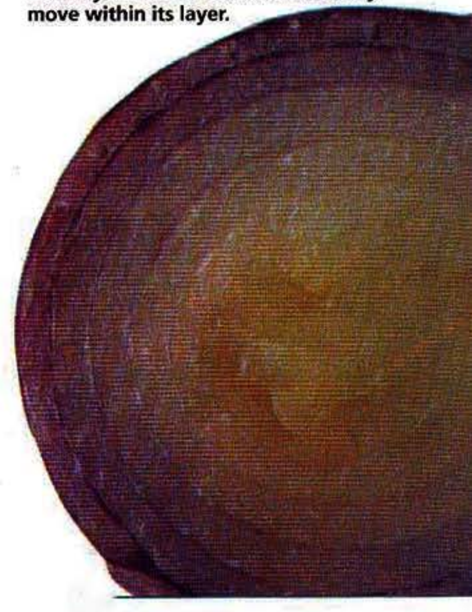


Fig. 5.6 A model of Rutherford's experiment. What does this experiment tell you about the size of the nucleus relative to the size of the atom?

The atom consists of a positively charged nucleus surrounded by negatively charged electrons. The nucleus is extremely tiny compared to the size of the atom.

The way in which electrons are arranged around the nucleus of an atom is very important because the *electron arrangement determines the chemical properties of the atom.*

The layers of an onion can be a simple model of how electrons occupy the space surrounding the nucleus. Imagine that electrons exist in layers around the nucleus. There are a fixed number of electrons in each layer and each electron can only move within its layer.



How are electrons arranged in the atom?

The electrons in an atom move around the nucleus in regions known as **electron shells** (Fig. 5.7). Each electron shell can only hold a certain number of electrons.

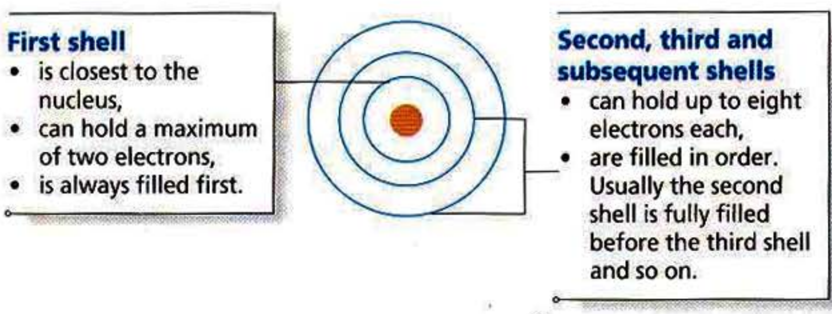


Fig. 5.7 Electron shells in an atom for elements with proton numbers 1 to 18

The electron arrangements of a hydrogen atom and a magnesium atom are shown in Fig. 5.8. *The way the electrons are arranged in an atom* is also called the atom's **electronic structure** or **electronic configuration**.

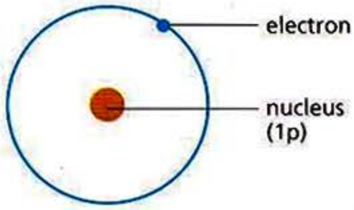


Fig. 5.8(a) **Electronic structure (electronic configuration) of a hydrogen atom**

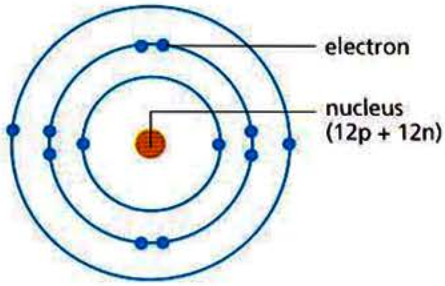


Fig. 5.8(b) **Electronic structure (electronic configuration) of a magnesium atom**

A magnesium atom (proton number 12) has two electrons in the first shell, eight electrons in the second shell and two electrons in the third shell. Thus, its electronic structure or electronic configuration can also be represented as (2, 8, 2).

Quick Check

An atom has an electronic configuration (2, 8, 4) and a nucleon number of 28. How many protons, neutrons and electrons are there in the atom?

What are valence electrons?

The *shell that is furthest from the nucleus* is called the **outer shell** or the **valence shell**. The electrons in this shell are called **valence electrons** (or valency electrons). A diagram of an atom's **outer electronic structure** *shows only the valence electrons in the valence shell.* An example is shown for magnesium (Fig. 5.9).

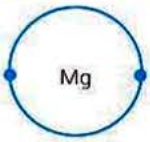


Fig. 5.9 **Outer electronic structure of a magnesium atom**

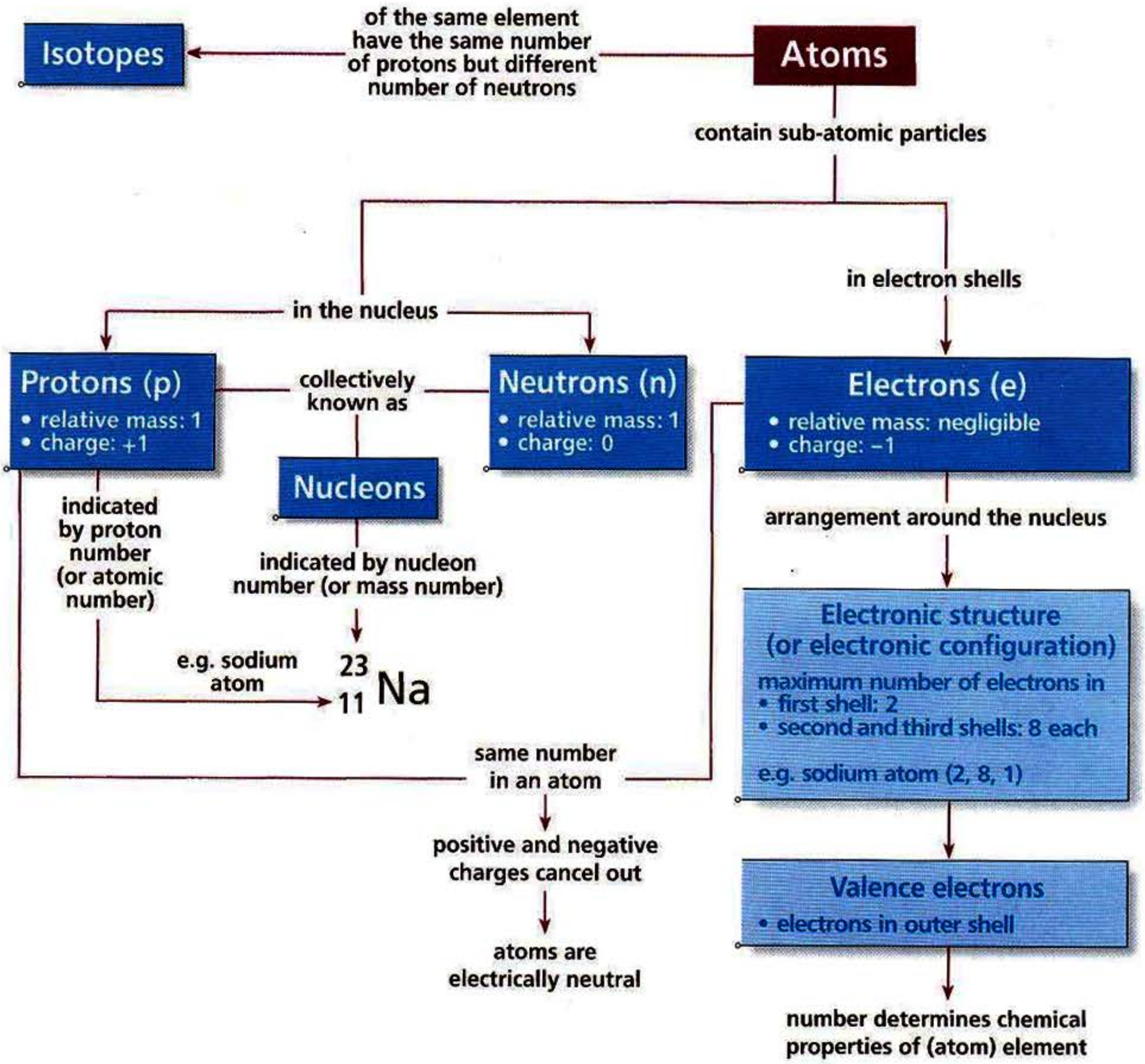
The chemical properties of an element depend on the number of valence electrons. For example, sodium (2, 8, 1) and potassium (2, 8, 8, 1) have similar chemical properties because their electronic configurations are similar. In each case, there is one electron in the outer shell.

The electronic configurations of 20 elements are shown in Fig. 5.10. It shows part of the Periodic Table which is a way of classifying elements. In the **Periodic Table**, *elements are arranged in order of increasing proton number.* There are *seven (horizontal) rows of elements* called **periods** and *eight (vertical) columns of elements* called **groups**.

Notice that elements that have the same number of valence electrons are in the same group. This means that elements in the same group have similar chemical properties. In the next chapter, we shall learn how the number of valence electrons determine the way different elements react with one another.

Both sodium and potassium react explosively with water to produce an alkali.

Concept Map


Atoms
of the same element have the same number of protons but different number of neutrons
Isotopes
contain sub-atomic particles
in the nucleus
in electron shells
Protons (p)
• relative mass: 1
• charge: +1
collectively known as
Neutrons (n)
• relative mass: 1
• charge: 0
Electrons (e)
• relative mass: negligible
• charge: −1
Nucleons
indicated by proton number (or atomic number)
indicated by nucleon number (or mass number)
e.g. sodium atom
$^{23}_{11}Na$
arrangement around the nucleus
Electronic structure (or electronic configuration)
maximum number of electrons in
• first shell: 2
• second and third shells: 8 each
e.g. sodium atom (2, 8, 1)
same number in an atom
positive and negative charges cancel out
atoms are electrically neutral
Valence electrons
• electrons in outer shell
number determines chemical properties of (atom) element

Exercise 5

Foundation

1. How many protons, neutrons and electrons are there in an atom of aluminium? (Nucleon number of Al = 27)

	Protons	Neutrons	Electrons
A	10	27	13
B	13	14	13
C	13	14	27
D	27	13	14

2. The diagram shows the nucleus of an atom of X. X is represented as ________.

represents a neutron

represents a proton

A $^{4}_{3}X$ **B** $^{7}_{3}X$

C $^{3}_{4}X$ **D** $^{7}_{4}X$

3. The table below gives the structure of five atoms.

Atom	Number of electrons	Number of neutrons	Number of protons
V	5	5	5
W	5	6	5
X	6	6	6
Y	7	7	7
Z	6	8	6

Which two pairs of atoms are isotopes?

	First pair	Second pair
A	V and W	X and Y
B	V and W	X and Z
C	W and X	Y and Z
D	W and Y	X and Z

4. Which pair of atoms contains the same number of neutrons?

A $^{114}_{48}Cd$ and $^{119}_{50}Sn$. **B** $^{59}_{27}Co$ and $^{59}_{28}Ni$.

C $^{133}_{55}Cs$ and $^{132}_{54}Xe$. **D** $^{63}_{29}Cu$ and $^{65}_{29}Cu$.

5. Copy and complete the table below to show the number of particles present in an atom of lithium and chlorine.

Element	Lithium	Chlorine
Proton number	3	17
Nucleon number		35
Electrons in one atom	3	
Protons in one atom		
Neutrons in one atom	4	

6. Gold is a precious metal. The symbol for gold is $^{197}_{79}Au$.

a) Complete the following sentence by writing in the missing words.

At the centre of an atom of gold is the ________ that contains 79 ________ and 118 ________.

b) Complete the following table by adding the missing information.

Particles in gold	Relative charge	Relative mass
Proton		
Neutron		
Electron		

7. Carbon-12 ($^{12}_{6}C$) and carbon-14 ($^{14}_{6}C$) are two isotopes of carbon. Carbon-14 is used for dating very old objects containing carbon such as wood.

   a) What is meant by the term 'isotopes'?

   b) Copy and complete the table below.

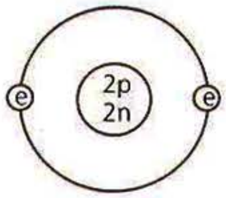
Per atom	Isotope	
	Carbon-12	Carbon-14
Number of protons		
Number of electrons		
Number of neutrons		

8. In what ways are the atoms $^{16}_{8}O$ and $^{17}_{8}O$

   a) similar to each other?

   b) different from each other?

9. In this question, 'e' represents an electron, 'n' a neutron and 'p' a proton.



   a) The diagram above shows an atom of helium.
      i) What is its proton number?
      ii) What is its mass number?
      iii) Draw a similar diagram to show the atomic structure of an isotope of this atom. The isotope has a nucleon number of 3.

   b) An incomplete diagram of the nucleus of an atom of another element is shown below.

      i) Complete the diagram to show the electronic arrangement of this atom.
      ii) Write the electronic configuration of the element.

Challenge

1. Natural silver consists of two types of atoms.

Type of atom	$^{107}_{47}Ag$	$^{109}_{47}Ag$
Relative percentage present	51.4%	
Protons in one atom		
Electrons in one atom		
Neutrons in one atom		

   a) Copy and complete the above table.

   b) What name is given to these two types of silver atoms?

2. Deuterium, an isotope of hydrogen, is represented in symbols as $^{2}_{1}D$. Another isotope of hydrogen is tritium (T).

   a) i) Tritium has two neutrons. Write the symbol for an atom of tritium, as has been written for deuterium.
      ii) Write the formula of the compound formed between tritium and oxygen.

> **Problem-solving strategy:**
> Isotopes of an element have similar chemical properties. Thus, the compound formed between tritium and oxygen is similar to the compound formed between hydrogen and oxygen.

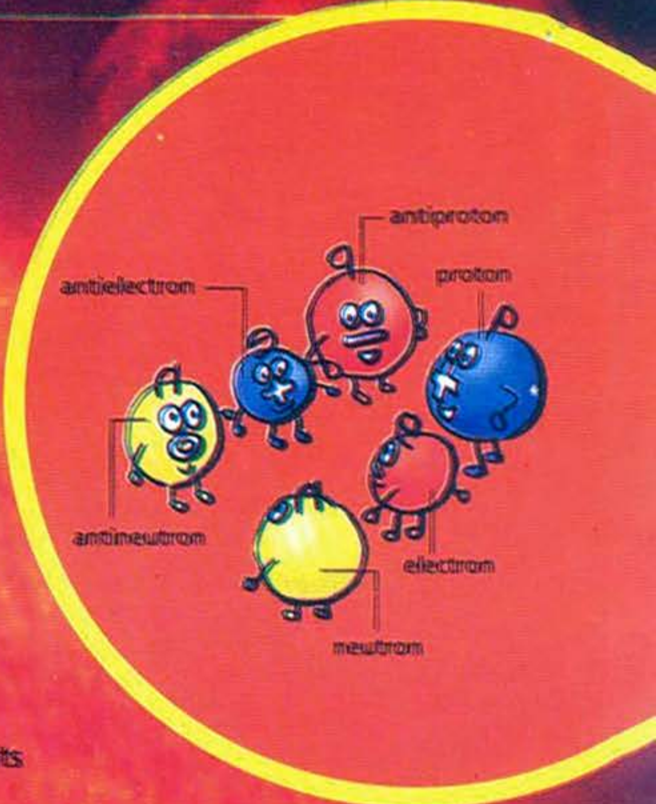
   b) Give the number of:
      i) electrons in HD.
      ii) protons in D_2.
      iii) neutrons in DT.

Chemistry Today

Matter is made up of atoms. In this chapter, we have seen that atoms are made of the sub-atomic particles protons, neutrons and electrons. For every particle that exists, there is an antiparticle. Antiparticles make up antimatter.

There are antielectrons called positrons — they are similar to electrons except that each positron has a positive charge (e^+). There are also antineutrons and antiprotons. An antiproton has the same mass as a proton but has a negative electric charge (p^-).

If matter and antimatter meet, they destroy each other and a large amount of energy is released. This is called annihilation. Annihilation occurs on the surface of the Sun, producing violent and fiery bursts of energy!



CRITICAL THINKING

1. An atom of $^{20}_{10}$Ne loses two protons and two neutrons from its nucleus and forms the atom *X*.
   i) What would be the number of protons and neutrons in the atom of *X* formed?
   ii) What is the symbol for *X*, including proton and nucleon numbers?
2. Explain why neutrons can be thought of as being made up of a proton combined with an electron.
3. An atom of $^{21}_{11}$Ne loses a proton from its nucleus and forms the atom *Y*.
   i) What would be the number of protons and neutrons in the atom *Y* formed?
   ii) What is the symbol for *Y*, including proton and nucleon numbers?

(Note: It has been shown that a neutron will decay into an electron, a proton and an antineutron.)

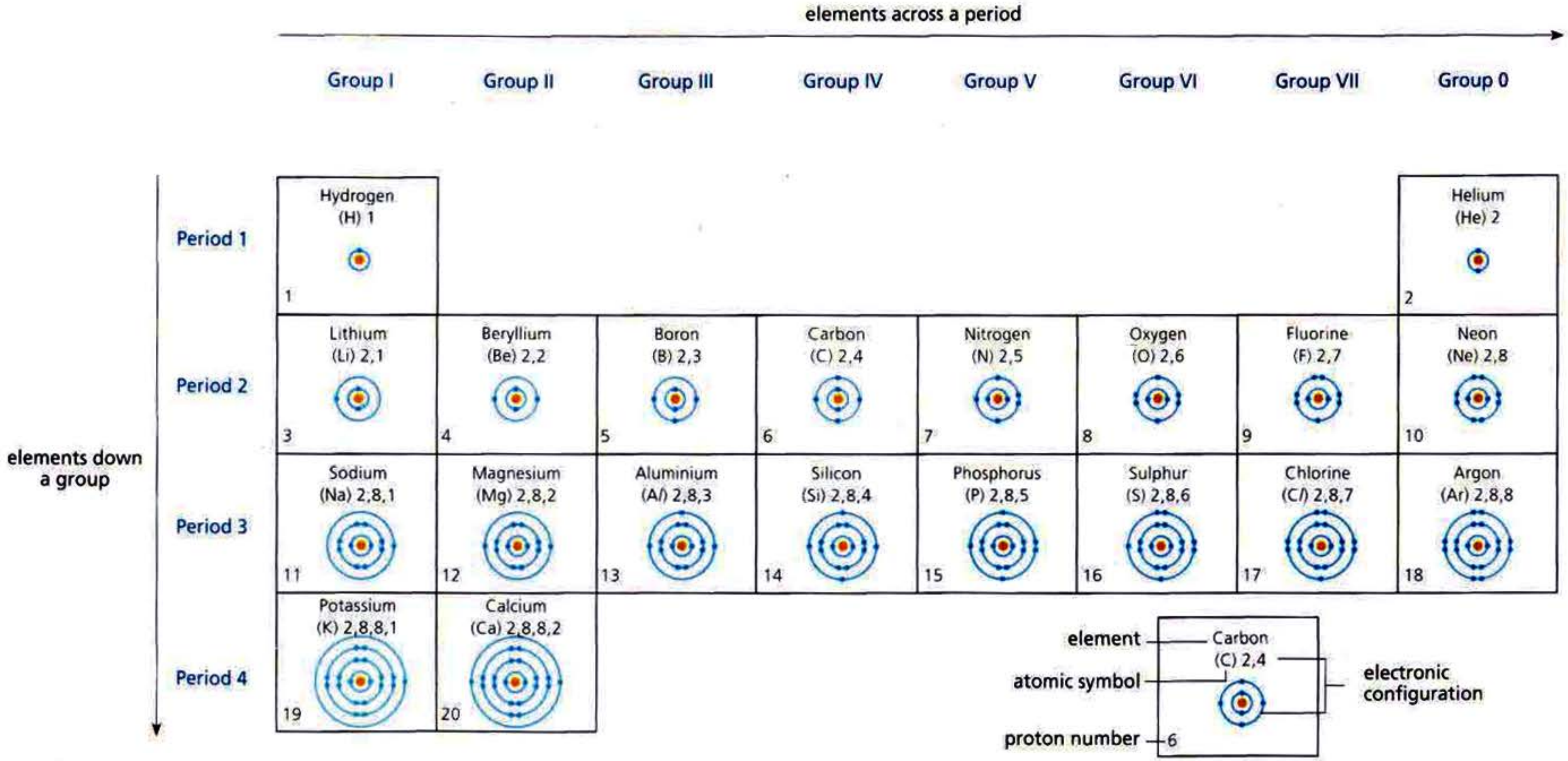


Fig. 5.10 The electronic configurations of elements with proton numbers 1 to 20 (The relative sizes of the atoms are not drawn to scale in this table.)

Key ideas

1. Isotopes are atoms of the same element with the same number of protons but different number of neutrons.
2. Isotopes have the same chemical properties but slightly different physical properties.
3. Electrons are arranged in electron shells.
4. The first electron shell can hold a maximum of two electrons. The second and third shells usually hold up to eight electrons.
5. The way in which electrons are arranged in electron shells is called the electronic structure or the electronic configuration of the atom.
6. Valence electrons are electrons found in the outer shell of an atom. They are important because they affect the chemical properties of an element.

Test Yourself 5.2

Worked Example

Bromine has two types of atoms $^{79}_{35}Br$ and $^{81}_{35}Br$.

a) What is the difference between the two atoms?
b) $^{19}_{9}F$, $^{35.5}_{17}Cl$ and $^{80}_{35}Br$ are in the same group of the Periodic Table. Using Fig. 5.10, deduce the number of valence electrons an atom of bromine has.

Thought Process

a) $^{79}_{35}Br$ and $^{81}_{35}Br$ are isotopes. They contain the same number of protons but different number of neutrons.
b) Look at the electronic configurations of fluorine and chlorine in Fig. 5.10. Both fluorine and chlorine have atoms that contain seven electrons in the outer shell. Observe also that elements in the same group have the same number of valence electrons. Since fluorine, chlorine and bromine are in the same group, bromine must have seven valence electrons.

Answer

a) $^{79}_{35}Br$ atoms have 44 neutrons in the nucleus and $^{81}_{35}Br$ atoms have 46 neutrons in the nucleus. $^{79}_{35}Br$ and $^{81}_{35}Br$ are two isotopes of bromine.
b) Seven

Questions

1. The symbol for aluminium is *Al*. Draw and write down the electronic configuration of the aluminium atom.

2. Draw the electronic structures of sodium (proton number 11) and lithium (proton number 3). Compare the structures and state one similarity and one difference between them.

Chapter 6

Ionic Bonding

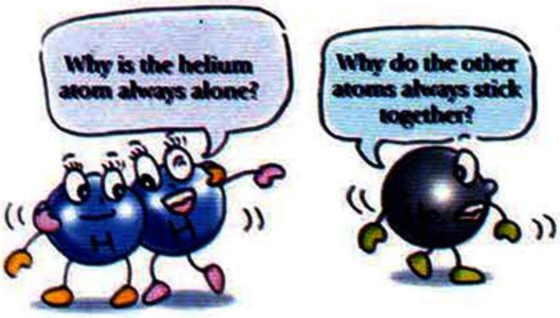
Neon-lit signs

Chapter Outline

The element neon is used as a gas in advertising signs. Neon is monatomic. Very few elements exist as individual atoms. Most of the matter in the world is made up of atoms that are chemically combined in some way. Why is that so? You will find out in this chapter.

6.1 The Stable Noble G

e element neon is used to produce the bright lights that you see on some advertising signs. In chapter 4, you learnt that neon (Ne) is monatomic. Neon belongs to a group of elements known as noble gases. The other noble gases are helium (He), argon (Ar), krypton (Kr), xenon (Xe) and radon (Rn).



In what way do noble gases behave differently from other elements?
Atoms of noble gases are **unreactive** or **stable**. This means that they do not combine with other atoms or form compounds. They exist as individual atoms.

What is the noble gas structure?
Examine the valence shells of the noble gases (Fig. 6.1).

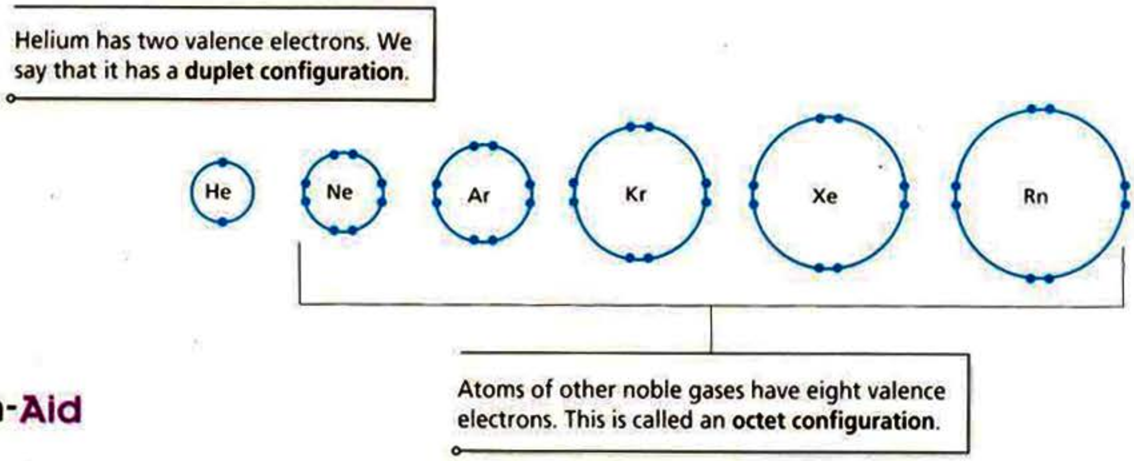


Fig. 6.1 Valence shells of noble gases

Chem-Aid

A valence electron is an electron found in the outer shell. It may also be called a valency electron, an outer electron or an outermost electron.

An atom is stable if it has a duplet or octet configuration. A duplet or octet configuration is also known as a noble gas structure or a **noble gas configuration**.

Atoms react in order to have the noble gas structure.

How do atoms achieve the noble gas structure?
Atoms of elements (besides the noble gases) react to achieve the noble gas structures. They do so by losing, gaining or sharing valence electrons. When atoms lose or gain electrons, they form ions.

Link

When atoms share electrons, molecules are formed. Read more about this type of bonding in chapter 7.

6.2 | Forming Ions

Atoms are electrically neutral because they have equal numbers of protons and electrons. An atom becomes an ion if it loses or gains electrons. In an **ion**, the number of protons is different from the number of electrons. Thus, an ion carries positive or negative charges.

An **ion** is *a charged particle formed from an atom or a group of atoms by the loss or gain of electrons.*

Link

Recall what you have learnt about atomic structure in chapter 5.

What type of ions do metals and non-metals form?

Examine Tables 6.1 and 6.2. What do you notice about the ions formed by metals and non-metals?

Name of metal	Name of ion	Formula of ion
sodium	sodium ion	Na^+
magnesium	magnesium ion	Mg^{2+}
aluminium	aluminium ion	Al^{3+}

Table 6.1 Ions of some metals

Name of metal	Name of ion	Formula of ion
chlorine	chloride ion	Cl^-
oxygen	oxide ion	O^{2-}
sulphur	sulphide ion	S^{2-}

Table 6.2 Ions of some non-metals

Metals form positively charged ions (**cations**) whereas non-metals form negatively charged ions (**anions**).

Chem-Aid

1. Remember: an electron has a negative charge.
2. 'Electronic configuration' is the same as 'electronic structure'.

How do atoms of metals become positive ions (cations)?

Atoms of metals tend to *lose valence electrons* to form *positive ions* or *cations*. In this way, they achieve a noble gas structure. For example, see the formation of a sodium ion in Fig. 6.2.

A sodium atom has 11 protons and 11 electrons. Its electronic configuration is (2, 8, 1).

To attain an octet configuration, the atom loses one valence electron. It forms a sodium ion.

The sodium ion that is formed has one positive charge because it has 11 protons but 10 electrons.

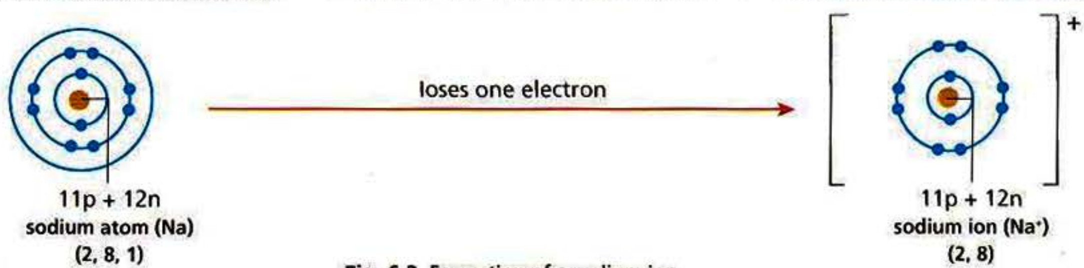


Fig. 6.2 Formation of a sodium ion

Quick Check

An ion contains 12 protons, 12 neutrons and 10 electrons. What is the charge on the ion? Name the ion.

A calcium atom has an electronic configuration of (2, 8, 8, 2). To attain an octet configuration, the atom loses two valence electrons. It forms a calcium ion (2, 8, 8). The calcium ion carries two positive charges (Fig. 6.3).

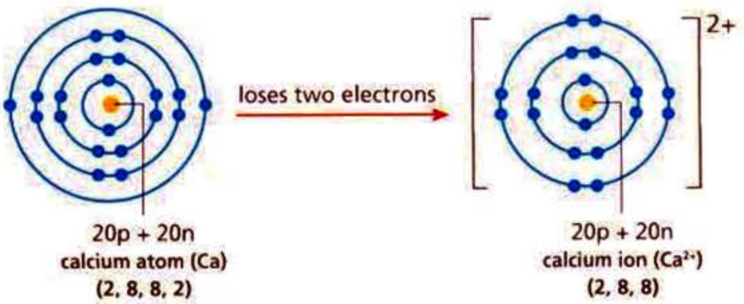


Fig. 6.3 Formation of a calcium ion

How do atoms of non-metals become negative ions (anions)?
Atoms of non-metals tend to *gain electrons* to form *negative ions* or *anions*. For example, a chlorine atom gains an electron to form a chloride ion (Fig. 6.4). Unlike the chlorine atom, the chloride ion has an octet configuration. The chloride ion carries one negative charge.

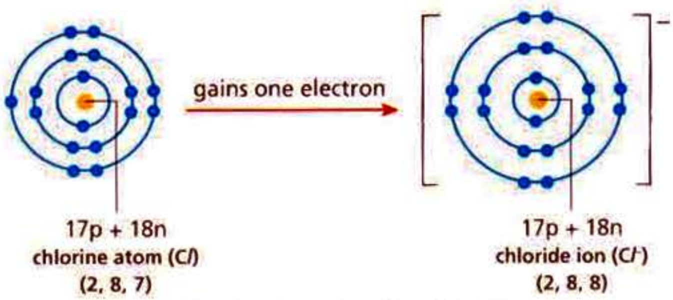


Fig. 6.4 Formation of a chloride ion

Fig. 6.5 shows the formation of an oxide ion from an oxygen atom.

The oxygen atom has an electronic configuration of (2, 6). It gains two electrons to form an oxide ion.

The oxide ion has an octet configuration. The oxide ion carries two negative charges because it has two more electrons than protons. Hence, it has the symbol O^{2-}.

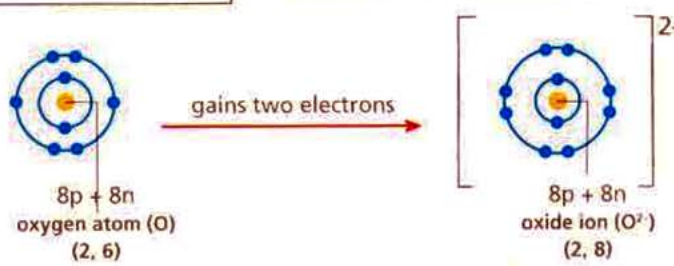


Fig. 6.5 Formation of an oxide ion

Common Ions and Their Charges

Tables 6.3(a) and 6.3(b) give the names and formulae of some cations and anions.

Charge on cation	Name of cation	Formula of cation
+1	ammonium	NH_4^+
	hydrogen	H^+
	potassium	K^+
	silver	Ag^+
	sodium	Na^+
+2	calcium	Ca^{2+}
	copper(II)	Cu^{2+}
	magnesium	Mg^{2+}
	zinc	Zn^{2+}
	iron(II)	Fe^{2+}
+3	iron(III)	Fe^{3+}
	aluminium	Al^{3+}

Table 6.3(a) The names and formulae of some cations

Cations are usually ions of metals. The ammonium ion (NH_4^+) and hydrogen ion (H^+) are some exceptions.

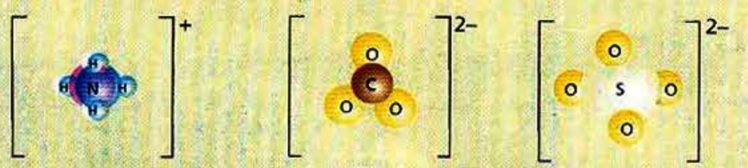
Some metals can form more than one ion. For example, iron forms Fe^{2+} and Fe^{3+}.

For these metals, the charge on the ion is shown in the name of the compound formed. For example, iron(III) chloride contains the Fe^{3+} ion.

Charge on cation	Name of cation	Formula of cation
−1	bromide	Br^-
	chloride	Cl^-
	hydroxide	OH^-
	nitrate	NO_3^-
−2	carbonate	CO_3^{2-}
	oxide	O^{2-}
	sulphate	SO_4^{2-}

Table 6.3(b) The names and formulae of some anions

Some ions are made up of groups of atoms. These ions are called **polyatomic ions**. Examples of polyatomic ions are the ammonium ion (NH_4^+), carbonate ion (CO_3^{2-}) and sulphate ion (SO_4^{2-}).



(a) An ammonium ion, NH_4^+ (b) A carbonate ion, CO_3^{2-} (c) A sulphate ion, SO_4^{2-}

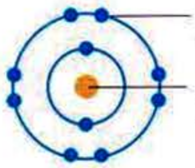
Key ideas

1. Noble gases are unreactive and do not form compounds because they have the duplet or octet configuration (noble gas structure).
2. An ion is a charged particle. In an ion, the number of protons is different from the number of electrons.
3. An atom forms an ion in order to achieve a noble gas structure.
4. Metals form positively charged ions (cations). Non-metals usually form negatively charged ions (anions).
5. A polyatomic ion is a group of atoms that carries a charge.

Test Yourself 6.1

Worked Example

What is the particle shown in the diagram on the right?



A F^- B N^{3-}
C Ne D O^{2-}

Thought Process

Refer to the Periodic Table on page vi. The element with eight protons is oxygen. Therefore, this must be an oxygen atom. There are ten electrons in the shells (2 + 8), therefore the ion must have a charge of 8 − 10 = −2. The ion is O^{2-}.

Answer

D

Questions

1. By referring to the Periodic Table, complete the table below to show the numbers of protons, neutrons and electrons in the ions H^+, Be^{2+}, F^-.

Ion	Number of protons	Number of neutrons	Number of electrons
H^+			
Be^{2+}			
F^-			

2. An ion contains seven protons, seven neutrons and ten electrons. By referring to the Periodic Table, write the formula of the ion.

3. What ions are present in copper(I) oxide and copper(II) sulphate?

6.3 | Forming Ionic Bonds

When metals react with non-metals, an ionic compound is formed. For example, sodium reacts with chlorine to form an **ionic compound** called sodium chloride.

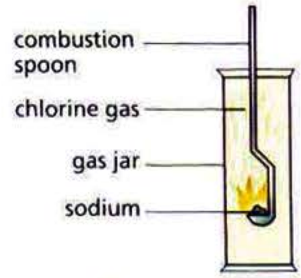


Fig. 6.6 The reaction between sodium and chlorine

What are the three steps involved in the formation of an ionic compound?
Take sodium chloride as an example.

1. **The formation of positive ions**
   Each sodium atom loses its single valence electron to form a positively charged sodium ion.

   $$\underset{(2,\,8,\,1)}{Na} \longrightarrow \underset{(2,\,8)}{Na^{+}} + e^{-}$$

   The sodium ion has the electronic configuration of the noble gas, neon.

Transfer of electron

2. **The formation of negative ions**
   Each chlorine atom gains an electron from a sodium atom to form a negatively charged chloride ion.

   $$\underset{(2,\,8,\,7)}{Cl} + e^{-} \longrightarrow \underset{(2,\,8,\,8)}{Cl^{-}}$$

   The chloride ion has the electronic configuration of the noble gas, argon.

3. **The formation of ionic bonds**
   Positive sodium ions and negative chloride ions are attracted to one another by **electrostatic attraction** to form sodium chloride.

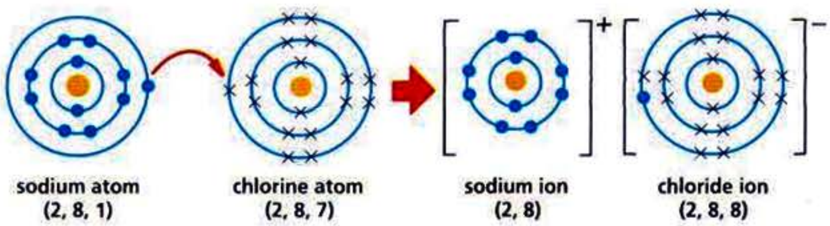
   $$Na^{+} + Cl^{-} \longrightarrow NaCl$$

Chem-Aid

1. Electrostatic attraction is the force that arises between positive and negative charges.
2. It is 'chloride ions', not 'chlorine ions'.

The electrostatic forces of attraction that hold the sodium ions and chloride ions together are called **ionic bonds**. An ionic bond may also be known as an **electrovalent bond**.

How do we show ionic bonding?
Fig. 6.7 shows two ways of representing the formation of an ionic bond in sodium chloride. They are called 'dot and cross' diagrams.



$$Na\bullet + {}^{\times\times}_{\times\times}\!\times Cl \longrightarrow [\,Na\,]^{+}\,[\,{}^{\times\times}_{\times\times}\!\bullet Cl\,]^{-}$$

Electrostatic attraction

Fig. 6.7 Two ways of showing the formation of an ionic bond in sodium chloride

In a 'dot and cross' diagram, dots represent the electrons of one atom, while crosses represent the electrons of another atom. In this example, the dots represent the electrons of the sodium atom, while the crosses represent the electrons of the chlorine atom.

Ionic bonds are formed when atoms of metals transfer their outer electrons to atoms of non-metals. Compounds that contain ionic bonds are called **ionic compounds**.

How is magnesium chloride formed?

Magnesium chloride is formed when magnesium reacts with chlorine. Fig. 6.8 shows how magnesium chloride is formed.

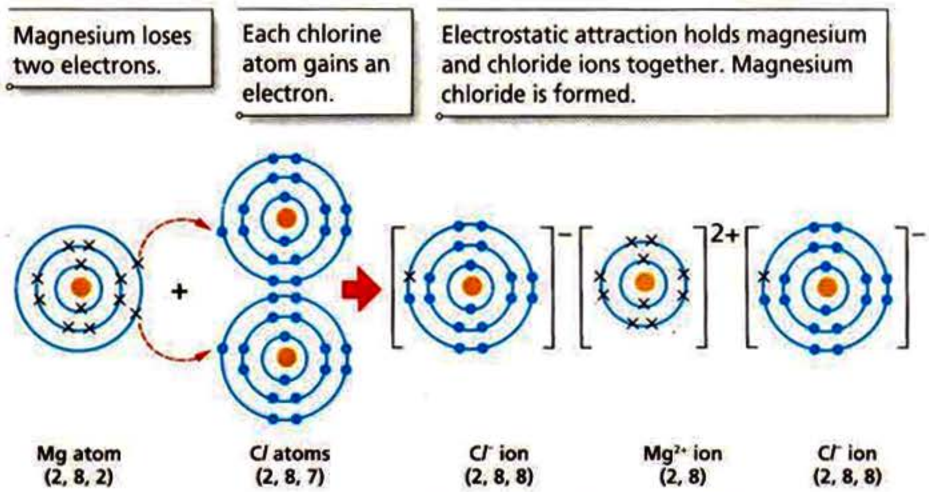


Fig. 6.8 Formation of ionic bonds in magnesium chloride

Chemical Formulae of Ionic Compounds

How do we determine the chemical formula of an ionic compound?

The formula of an ionic compound is constructed by balancing the charges on the positive ions with those on the negative ions. *All the positive charges must equal all the negative charges in an ionic compound.*

Consider the ionic compound, calcium chloride. The ions present are Ca^{2+} and Cl^-. Here, Ca^{2+} has two positive charges but Cl^- has only a single negative charge. Therefore, there must be two Cl^- to balance out the positive charges. The chemical formula is $CaCl_2$.

Example 1
What is the chemical formula of magnesium oxide?

Solution:
Magnesium forms Mg^{2+} and oxygen forms O^{2-} in the compound.
(Number of Mg^{2+} × positive charge on Mg^{2+}) balance out
(number of O^{2-} × negative charge on O^{2-})
1 × (+2) balance out 1 × (–2)

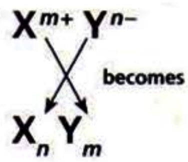
Therefore, the formula is MgO.

Example 2
What is the chemical formula of copper(II) hydroxide?

Solution:
Copper forms Cu^{2+} and hydroxide ions are OH^-.
1 × (+2) balance out 2 × (–1)

The formula is $Cu(OH)_2$. Notice that polyatomic ions need to be enclosed within brackets if the subscript is two and above.

To derive the formula, you could also
- write down the ions with the charges, e.g. X^{m+} Y^{n-},
- move the values m and n diagonally (but without the charges). Formula is X_nY_m.



Example 3
What is the chemical formula of aluminium oxide?

Solution:
The ions are Al^{3+} and O^{2-}, so its formula is Al_2O_3.

$$\overset{3+}{Al} \quad \overset{2-}{O} \longrightarrow Al_2O_3$$

Example 4
What is the chemical formula of calcium carbonate?

Solution:
'2' is a common factor of the subscripts of Ca and CO_3, so it is removed. As a result, no brackets are needed for CO_3.

$$\overset{2+}{Ca} \quad \overset{2-}{CO_3} \longrightarrow Ca_2(CO_3)_2 \longrightarrow CaCO_3$$

Chem-Aid

General rules for writing chemical formulae:

1. In naming ionic compounds, the symbol of the metal is usually written first.
2. The subscripts are the simplest set of whole numbers, e.g. MgO and not Mg_2O_2. There are exceptions to this rule, e.g. hydrogen peroxide, H_2O_2, and dinitrogen tetroxide, N_2O_4.
3. The subscript '1' is not written, e.g. NaCl and not Na_1Cl_1.
4. To indicate more than one of a polyatomic ion in the formula, place the ion within brackets, e.g. $Cu(OH)_2$ and not $CuOH_2$.

Key ideas

1. Metals react with non-metals to form ionic compounds.
2. Each ion in an ionic compound has the electronic configuration of a noble gas.
3. An ionic bond is defined as the electrostatic force of attraction between a positive ion and a negative ion.
4. An ionic bond is formed when electrons are transferred from a metallic atom to a non-metallic atom.
5. In an ionic compound, the total positive charge is equal to the total negative charge.

Test Yourself 6.2

Worked Example

What is the chemical formula of the compound formed when $^{24}_{12}X$ combines with $^{19}_{9}Y$?

A XY B X_2Y C XY_2 D X_2Y_2

Thought Process

- Since X has 12 protons, it must have 12 electrons arranged as (2, 8, 2). Similarly, Y must have nine electrons arranged as (2, 7).
- In order to attain noble gas structures, X needs to lose two electrons and Y needs to gain one electron. X will form the ion X^{2+} and Y will form the ion Y^{-}. Hence, the chemical formula of the compound is XY_2.

$^{24}_{12}\mathbf{X}$

X has a proton number of 12. Its atom has 12 electrons.

$\mathbf{X^{2+}\ Y^{1-}}$

$\mathbf{XY_2}$

Answer

C

Questions

1. $\boxed{\text{bromine, carbon, iron, sodium, sulphur}}$

   From the above list, select **four** pairs of elements that form ionic bonds when they react. What are the chemical formulae of these four compounds?

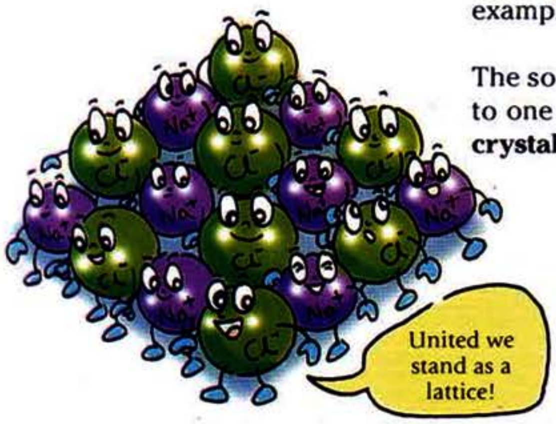
2. Refer to Table 6.3 and write the chemical formulae of the following compounds: sodium oxide, aluminium nitrate, calcium hydroxide, potassium sulphate and calcium sulphate.

Structure of Ionic Compounds

Ionic compounds form giant ionic structures

Ionic compounds form **giant ionic structures** similar to the one shown in Fig. 6.9. Sodium chloride (common salt) is a typical example of an ionic compound.

The sodium ions and the chloride ions are very strongly attracted to one another. They are arranged in a **giant lattice structure** or **crystal lattice**.



In a lattice, there are millions of sodium and chloride ions arranged in an orderly manner. These ions are held in place by ionic bonds throughout the entire lattice.

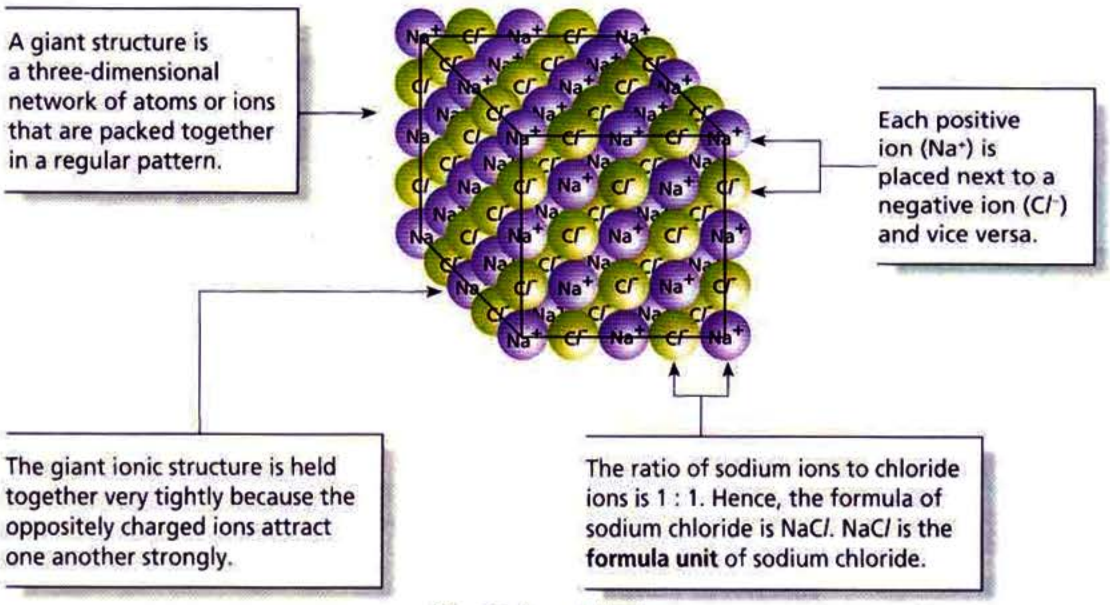


Fig. 6.9 A crystal lattice

Physical Properties of Ionic Compounds

Volatility — melting and boiling points

The lattice of an ionic compound is held together by strong ionic bonds between the ions. A large amount of energy is needed to overcome these strong bonds and to change an ionic compound from solid to liquid state. As a result, ionic compounds are solids at room temperature and pressure. We say ionic compounds are *non-volatile* substances. In fact, *many ionic compounds have high melting points* and *high boiling points*.

Solubility

Table 6.4 shows the solubility of some ionic compounds in water and in oil. Can you deduce a general rule for the solubility of ionic compounds in water and in oil?

Ionic compound	Solubility in water	Solubility in oil
calcium chloride	soluble	insoluble
potassium nitrate	soluble	insoluble
magnesium sulphate	soluble	insoluble

Table 6.4 Solubilities of some ionic compounds in water and in oil

Ionic compounds are usually soluble in water. This is because water molecules can separate the positive ions from the negative ions, causing them to dissolve (Fig. 6.10).

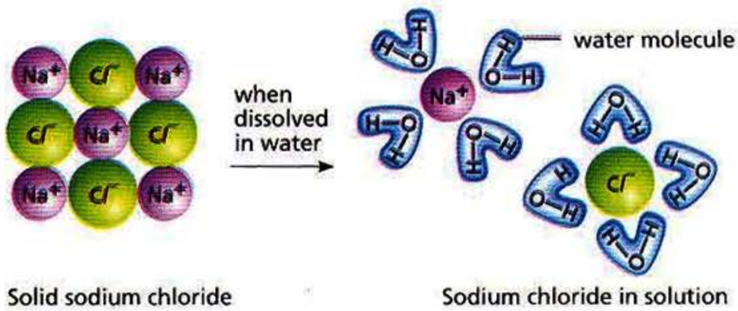


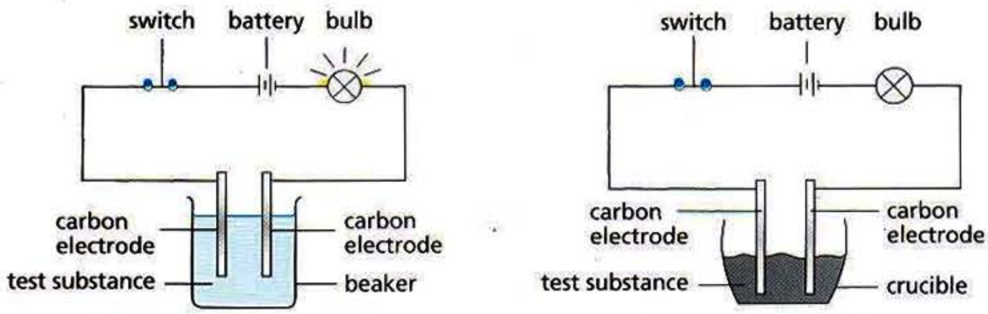
Fig. 6.10 The arrangement of Na^+ and Cl^- ions in the solid state and in solution

However, some ionic solids such as silver chloride and barium sulphate are insoluble in water.

Ionic compounds are insoluble in organic solvents. Organic solvents are compounds such as ethanol, petrol and turpentine.

Electrical conductivity

The circuits shown in Fig. 6.11 can be used to determine whether a substance can conduct electricity. If the substance conducts electricity, the bulb will light up when the switch is turned on. If the substance is a non-conductor of electricity, the bulb will not light up when the switch is turned on.



(a) Test substance is a liquid. **(b) Test substance is a solid.**

Fig. 6.11 Apparatus for testing electrical conductivity

Chem-Aid

1. Aqueous sodium chloride refers to a solution of sodium chloride in water.
2. Molten sodium chloride refers to sodium chloride that has been heated until it has completely melted (i.e. it is completely liquid).

For example, when *solid* sodium chloride is used as the test substance, the bulb does not light up. After *water is added* to dissolve the solid to form aqueous solution of sodium chloride, the bulb lights up. Similarly, when *molten* sodium chloride is used, the bulb lights up.

What does this experiment tell you?

Ionic compounds do not conduct electricity in the solid state because the ions are not free to move about. However, when an ionic compound is melted or dissolved in water to form an aqueous solution, it can conduct electricity. This is because the ions are free to move in the molten state or in aqueous solution.

Test Yourself 6.3

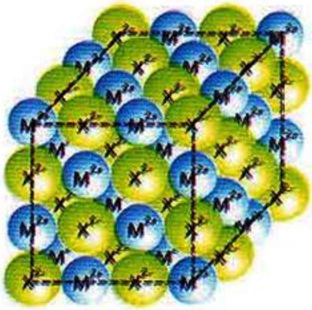
Questions

1. The properties of strontium chloride, $SrCl_2$, are as follows:
   - Melts at 875 °C
   - Dissolves readily in water

   a) Why does strontium chloride have a high melting point?
   b) Predict **two** other physical properties of strontium chloride.

2. Suggest one reason why magnesium oxide is used to line the inner surface of a furnace.

3. The crystal lattice of an ionic compound is shown below. Briefly describe how you deduce the chemical formula of this ionic compound.



Key ideas

1. An ionic compound consists of ions arranged in a giant ionic structure called a crystal lattice.
2. The physical properties of an ionic compound are as follows:
   - Non-volatile, high melting point, high boiling point
   - Usually soluble in water but insoluble in organic solvents
   - Good electrical conductor in the molten state or in aqueous solution but not in the solid state

Concept Map

Metals → **Attain noble gas configuration** by losing electrons to form → **Positive ions (cations)**

Non-metals → **Attain noble gas configuration** by gaining electrons to form → **Negative ions (anions)**

Positive ions (cations) and Negative ions (anions) are attracted to one another by electrostatic forces to form → **Ionic bonds**

Ionic bonds found in → **Ionic compounds**

Ionic compounds exist as → **Giant ionic structures (crystal lattice)**

Ionic compounds → **Properties:**
- contain ions
- high melting and boiling points
- non-volatile
- usually soluble in water but insoluble in organic solvents
- conduct electricity when molten or when dissolved in water

Exercise 6

Foundation

1. Which atom loses two electrons from its valence shell to form an ion?

   A Calcium B Carbon
   C Chlorine D Oxygen

2. Magnesium atoms have the electronic configuration (2, 8, 2). Oxygen atoms have the electronic configuration (2, 6). Which statement is true about the formation of magnesium oxide?

   A A magnesium atom accepts a pair of electrons from an oxygen atom.
   B A magnesium atom gives a pair of electrons to an oxygen atom.
   C An oxygen atom gives a pair of electrons to a magnesium atom.
   D An oxygen atom shares a pair of electrons with a magnesium atom.

3. Some properties of four substances are tabulated below. Which substance is likely to be sodium chloride?

Substance	Melting point (°C)	Ability to conduct electricity	
		when molten	in aqueous solution
A	–100	none	none
B	–85	none	good
C	115	none	(insoluble)
D	808	good	good

4. Talc is a mineral and has the formula $Mg_3Si_4O_{10}(OH)_2$. What is the charge on Si_4O_{10}?

   A 2– B 2+
   C 3+ D 4–

5. Write the symbols of **three** particles (atoms or ions) that have only two electrons. You may refer to the Periodic Table.

6. Why are ionic compounds **not** used to make perfumes?

7. *Y* is the ion of an element *X*. *Y* contains 13 protons, 14 neutrons and 10 electrons.

   a) What is the nucleon number of *Y*?
   b) Draw a 'dot and cross' diagram to show an atom of *X*.
   c) Predict the formula of the compound that contains *Y* and the oxide ion.

Challenge

1. Which statement about the formation of ionic bonds in the following compounds is correct?

	Compound	Formation of ionic bonds
A	$NaCl$	two electrons transferred
B	$CaCl_2$	three electrons transferred
C	MgO	four electrons transferred
D	Al_2O_3	six electrons transferred

2. Chlorine reacts with sodium to form the compound $NaCl$. Chlorine also reacts with phosphorus to form the compound PCl_3. What will be the chemical formula of the compound formed between sodium and phosphorus?

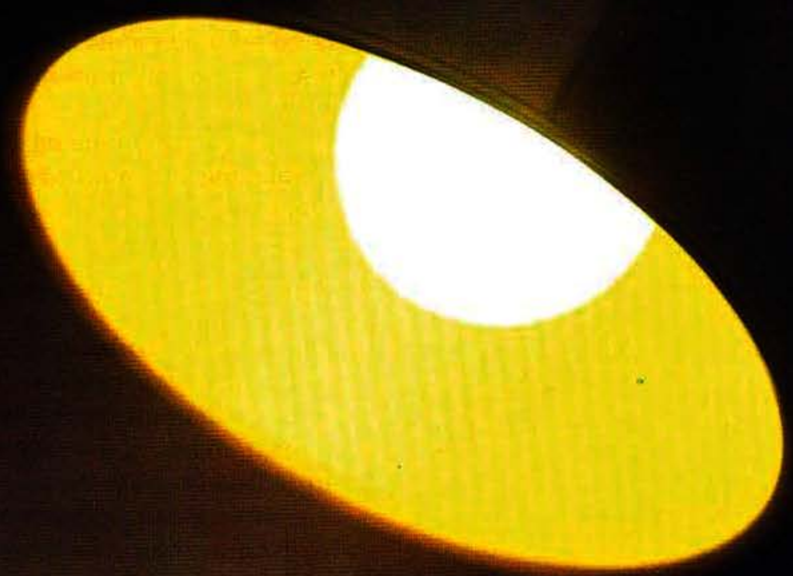
   A Na_2P_3 B Na_3P
   C NaP D NaP_3

3. Element *X* reacts with bromine to form an ionic compound, XBr_2.

   a) State, with reasons,
      i) the physical state of XBr_2 at room temperature.
      ii) whether *X* is a metal or a non-metal.
   b) Predict **two** other properties of XBr_2.
   c) Deduce the charge of ion *X* in the compound XBr_2.
   d) What is the formula of the oxide formed when *X* reacts with oxygen?

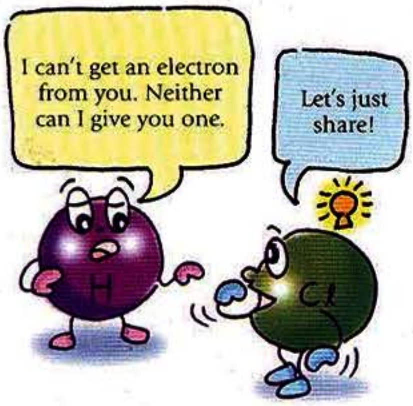
Chapter 7

Covalent and Metallic Bonding



Chapter Outline

Elements such as chlorine and tungsten are used to make a modern light bulb. Both these elements are stable without forming ionic compounds. How is this possible? What kind of chemical bonding exists in such elements? Find out in this chapter.



Sharing electrons

7.1 | Covalent Bonds

In the previous chapter, you learnt that non-metals react with metals to attain a stable noble gas structure. Non-metals can also attain such structures by reacting with one another.

How do non-metals react with one another?
Non-metals react with one another when their atoms share valence electrons. *The bond formed between atoms that share electrons* is called a **covalent bond**. After bonding, each atom attains the electronic configuration of a noble gas.

When atoms combine by sharing electrons, molecules are formed. A **molecule** is *a group of two or more atoms held together by covalent bonds*.

Molecules of Elements

How are electrons arranged in molecules of elements?
In chapter 4, you learnt that elements such as hydrogen and oxygen exist as diatomic molecules (molecules made up of two atoms). In fact, many non-metallic elements exist as molecules made up of two or more identical atoms. Let us look at how electrons are shared and arranged within molecules of some elements: hydrogen, chlorine and oxygen.

Hydrogen
A 'dot and cross' diagram of the hydrogen molecule is shown in Fig. 7.1. In this diagram, the electron of one hydrogen atom is represented by a '•'. '×' represents the electron of the other hydrogen atom.

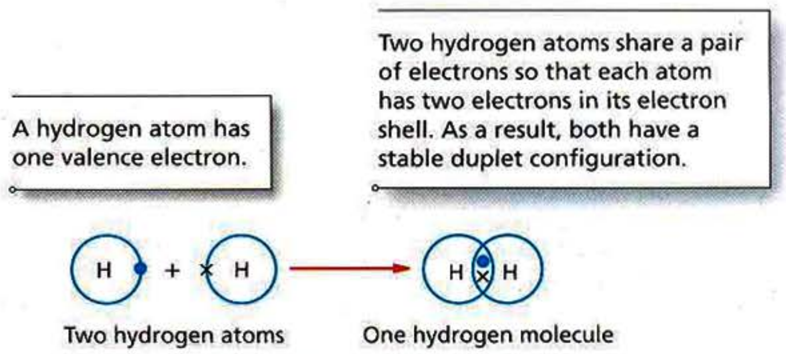


Fig. 7.1 Electron sharing in a hydrogen molecule

The *sharing of two electrons* forms a **single covalent bond**. A single covalent bond or a single bond is represented by a single line in the structural formula.

Table 7.1 shows the various ways of representing a hydrogen molecule.

'Dot and cross' diagram	Structural formula	Molecular formula	Model
H $\overset{\times}{\bullet}$ H	H—H	H_2	

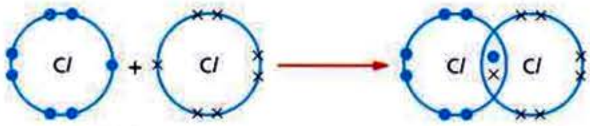
Table 7.1 **Different ways of representing a hydrogen molecule**

Chem-Aid

A structural formula shows how atoms are arranged in a molecule (see chapter 21).

Chlorine

A chlorine atom has seven valence electrons. It needs one more electron to form a stable octet structure of eight electrons. To attain this structure, two chlorine atoms combine to share a pair of electrons through covalent bonding as shown in Fig. 7.2.



Two chlorine atoms	One chlorine molecule

Fig. 7.2 **Electron sharing in a chlorine molecule. For simplicity, only valence electrons are shown.**

Quick Check

Do noble gases exist as diatomic molecules?

Table 7.2 shows the various ways of representing a chlorine molecule.

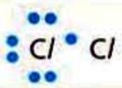
'Dot and cross' diagram	Structural formula	Molecular formula	Model
	Cl—Cl	Cl_2	

Table 7.2 **Different ways of representing a chlorine molecule**

Quick Check

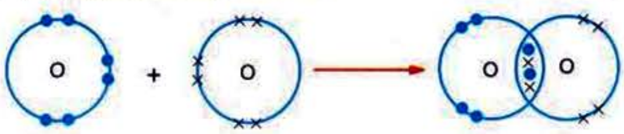
How many covalent bonds are there in one molecule of chlorine gas?

Oxygen

Fig. 7.3 shows the formation of an oxygen molecule.

An oxygen atom has only six valence electrons. It needs two more electrons to form a stable octet structure. Thus two oxygen atoms combine to form a molecule.

Each oxygen atom shares two of its electrons with another oxygen atom to form the covalent bond shown.



Two oxygen atoms	An oxygen molecule

Fig. 7.3 **Electron sharing in an oxygen molecule. A double covalent bond is formed.**

As you can see in Fig. 7.3, four electrons are shared between two oxygen atoms. Such a bond is called **a double covalent bond** or **double bond**. A double bond is formed when *two pairs of electrons are shared between two atoms*. A double bond is represented by '═' in a structural formula. Table 7.3 shows the different ways of representing the oxygen molecule.

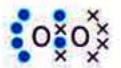
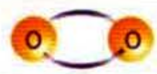
'Dot and cross' diagram	Structural formula	Molecular formula	Model
	O═O	O_2	

Table 7.3 Different ways of representing an oxygen molecule

Molecules of Compounds

How are electrons arranged in molecular compounds?

Molecules made from two or more different types of atoms linked together by covalent bonding are called **molecular compounds** or **covalent compounds**. Water, methane and carbon dioxide are examples of covalent compounds.

Water

Water is formed by the reaction of hydrogen with oxygen such that all three atoms attain noble gas configurations. Each water molecule contains two single covalent bonds.

Chem-Aid

1. Do note that a double covalent bond is different from two single covalent bonds.
2. The phrase 'noble gas configuration' can be used to describe either a duplet or an octet configuration.

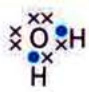
'Dot and cross' diagram	Structural formula	Molecular formula	Model
	O — H | H	H_2O	

Table 7.4 Different ways of representing a water molecule

Methane

Methane contains two elements, carbon and hydrogen. In a methane molecule, the carbon atom has an octet configuration while each hydrogen atom has a duplet configuration. The molecular formula of methane is CH_4. Methane has four single covalent bonds.

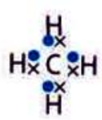
'Dot and cross' diagram	Structural formula	Molecular formula	Model
	H | H — C — H | H	CH_4	

Table 7.5 Different ways of representing a methane molecule

Carbon dioxide

Carbon dioxide is formed when carbon reacts with oxygen. The carbon atom must share two electrons each with two oxygen atoms in order to achieve a stable octet of electrons. Each oxygen atom shares two electrons and attains an octet configuration. Consequently, a carbon dioxide molecule contains two double covalent bonds. The molecular formula for carbon dioxide is CO_2.

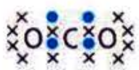
'Dot and cross' diagram	Structural formula	Molecular formula	Model
	O═C═O	CO_2	

Table 7.6 Different ways of representing a carbon dioxide molecule

Key ideas

1. **A covalent bond is the bond formed by the sharing of electrons between atoms.**
2. **A single covalent bond consists of one shared pair of electrons. A double covalent bond consists of two shared pairs of electrons.**
3. **A molecule is formed when two or more atoms are joined together by covalent bonds.**
4. **In a molecule, each atom has the electronic configuration of a noble gas.**

Test Yourself 2.1

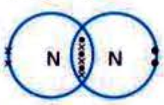
Worked Example

Draw the structure of a molecule of nitrogen. Only valence shells need to be shown.

Thought Process

The electronic configuration of a nitrogen atom is (2, 5). When two nitrogen atoms join together to form a molecule, a **triple bond** is formed. A triple bond contains *three* shared pairs of electrons.

Answer



A nitrogen molecule

Questions

1. Identify **four** covalent compounds from the following list:

$MgSO_4$, NH_3, $CHCl_3$, Br_2, $CaCl_2$, SO_2

2. Draw 'dot and cross' diagrams to show how different types of chemical bonds are formed when fluorine reacts with
   a) hydrogen. b) potassium.

3. Silicon reacts with chlorine to form the covalent compound $SiCl_4$.
   a) What is a covalent bond?
   b) Draw a 'dot and cross' diagram (showing only valence electrons) to show the structure of this compound.

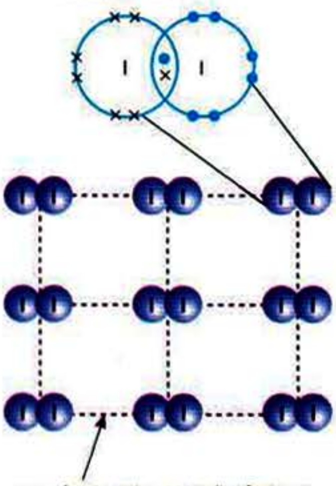


Fig. 7.4 Structure of solid iodine

7.2 | Structures and Properties of Covalent Substances

Covalent substances may exist as simple molecules or as giant structures.

Simple Molecular Structures

Iodine

Solid iodine is purplish-black in colour. When heated, it sublimes to form a purple gas. The purple gas is made up of diatomic molecules of iodine, I_2.

The arrangement of iodine molecules in solid iodine is shown in Fig. 7.4. *Within* each iodine molecule, the iodine atoms are held together by *strong* covalent bonds. *Between* the iodine molecules in the solid, there are only *weak* **van der Waals' forces** holding the molecules together. These weak forces break down on gentle heating. This is why solid iodine sublimes on heating.

Since the iodine molecules are loosely held together, they exist as individual or discrete molecules. We can say that iodine exists as simple discrete covalent molecules. We can also say that iodine has a **simple molecular structure**.

Chem-Aid

Van der Waals' forces are also known as intermolecular forces.

Methane

In a molecule of methane, CH_4, the four C–H covalent bonds are strong. However, weak van der Waals' forces *between* methane molecules hold them together loosely. Therefore, methane exists as a gas at room temperature and pressure (Fig. 7.5).

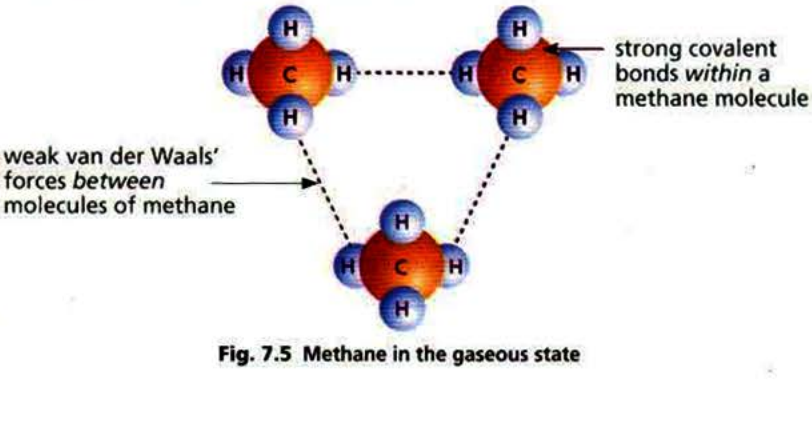


Fig. 7.5 Methane in the gaseous state

What are the physical properties of simple molecular covalent compounds?

Many covalent substances have simple molecular structures. The physical properties of these substances are described below.

Volatility — melting and boiling points

The melting and boiling points of some covalent substances are shown in Table 7.7.

Covalent substance	Melting point (°C)	Boiling point (°C)
carbon dioxide	–56	–79
chlorine	–101	–35
hydrogen	–259	–253
methane	–183	–161
oxygen	–214	–183
water	0	100

Table 7.7 Melting and boiling points of some covalent substances

Many simple covalent substances, especially those made up of small molecules, are liquids or gases at room temperature. They are volatile or have high volatility (they evaporate easily). This occurs because the forces between the molecules are weak compared to the covalent bonds within the molecules. (See Fig. 7.6 for the example on water.) Very little heat energy is required to overcome the intermolecular forces. Thus, these substances have low melting points and boiling points.

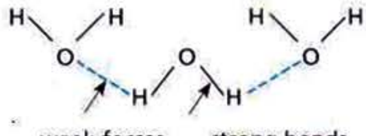


Fig. 7.6 Different types of bonding exist *between* and *within* water molecules

Simple covalent substances made up of larger molecules are solids at room temperature. This is because the intermolecular forces of attraction are stronger between larger molecules.

Chem-Aid

When liquid bromine evaporates, the bromine molecules do not break up into atoms, but merely move apart (Fig. 7.7).

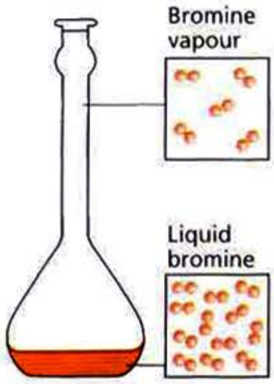


Fig. 7.7 Bromine and its structures in the gaseous state and liquid state at room temperature

Solubility

Most covalent molecules are *insoluble in water and soluble in organic solvents*. However, there are some exceptions, for example, alcohol and sugar are covalent compounds that are soluble in water. Some covalent molecules (e.g. chlorine and hydrogen chloride) **dissociate** when they dissolve in water. You will learn more about dissociation in chapter 11.

Electrical conductivity

Most covalent elements or compounds do not conduct electricity whether in the solid, liquid or gaseous state. This is because they do not have free-moving ions or electrons to conduct electricity. There are exceptions. Carbon, in the form of graphite, conducts electricity. Hydrogen chloride, sulphur dioxide and ammonia react with water to form solutions that conduct electricity.

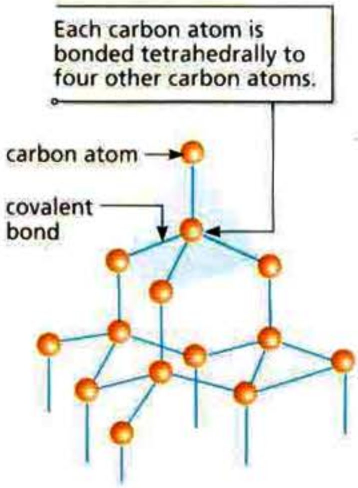


Fig. 7.8 Giant molecular structure of diamond

Giant Molecular Structures

Diamond

Diamond is one of the allotropes of carbon. ***Allotropes*** are *different forms of the same element.*

In diamond, each carbon atom is joined to four other carbon atoms by strong covalent bonds (Fig. 7.8). Diamond has a giant molecular structure.

What are the physical properties of diamond?

Diamond is a very hard substance. It is not easily scratched or worn out.

Diamond exists as a solid in nature. It has high melting and boiling points. In fact, the melting point of diamond is 3500°C and its boiling point is 4800 °C. Diamond does not conduct electricity. It is insoluble in water.

How does the structure of diamond account for its physical properties?

Diamond has all the typical properties of a giant covalent molecule. A crystal of diamond contains millions of carbon atoms joined by strong covalent bonds. A lot of energy is required to break these strong covalent bonds. This explains why diamond is so hard and difficult to melt.

In the diamond structure, all the valence electrons of the carbon atoms are used for bonding. There are therefore no free electrons that move through the structure. Hence, diamond *cannot* conduct electricity.

Quick Check

Suggest what might be used to cut diamond to produce gemstones.

What are the uses of diamond?

Diamonds are rare and very precious. They are used as gemstones in jewellery. However, the main uses of diamond depend on its hardness as well as high melting and boiling points. Synthetic diamonds produced under high pressures and temperatures are used at the tips of drills and other cutting tools. They are used for drilling, grinding and polishing very hard surfaces.

Graphite

Graphite is another allotrope of carbon. It is made of layers of carbon atoms. Within each layer, each carbon atom forms strong covalent bonds with three other carbon atoms, indicated by 1, 2 and 3 in Fig. 7.9. This forms rings of six carbon atoms that are joined together to form two-dimensional flat layers. Therefore, each layer is a giant molecule. The layers of carbon atoms, which are held by weak van der Waals' forces, lie on top of each other.

Pencil lead is a mixture of graphite and clay.

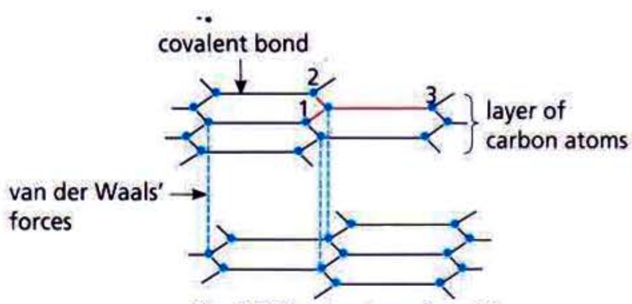


Fig. 7.9 The structure of graphite

How does the structure of graphite explain its physical properties?

In graphite, the bonds *within* each layer are strong and difficult to break. Like diamond, graphite is a solid. It has high melting and boiling points.

However, the forces of attraction *between* the layers of carbon are very weak. The layers can slide over each other. This explains why graphite is soft and slippery.

In graphite, each carbon atom has one outer shell electron that is not used to form covalent bonds (Fig. 7.10). These electrons are delocalised, that is, they can move along the layers from one carbon atom to the next when graphite is connected to a battery. Hence, graphite is a good conductor of electricity.

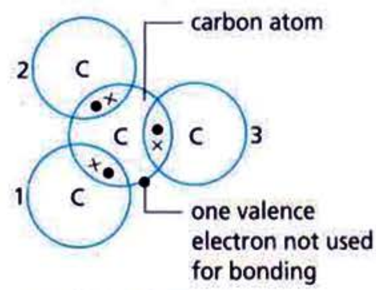


Fig. 7.10 Covalent bonding in graphite

What are the uses of graphite?

Table 7.8 shows how the uses of graphite are dependent on its physical properties.

Property of graphite	Uses of graphite that depend on this property
• slippery • does not decompose even at very high temperatures • does not attack rubber	Graphite is used as a dry lubricant. Oil attacks rubber, but graphite does not. It is used to lubricate machine parts that contain rubber.
• soft (its layers of carbon atoms can be rubbed off easily)	Graphite is baked with clay and made into pencil lead. Since it is soft, the carbon layers flake off and stick to paper when we write.
• good conductor of electricity • fairly unreactive	Graphite is used as brushes for electric motors. It is also used to make inert electrodes for electrolysis. You will learn more about this use in chapter 15.

Table 7.8 Properties and uses of graphite

Silicon (IV) oxide

Sand is actually silicon(IV) oxide, also commonly known as silica. Silicon(IV) oxide has a **giant molecular structure**. This is because all the atoms of silicon and oxygen join together to form a three-dimensional structure.

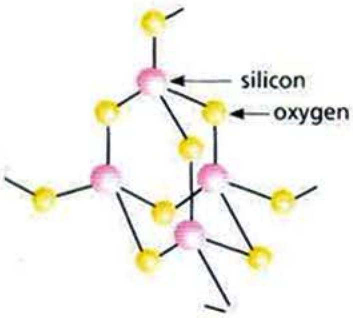


Fig. 7.11 Covalent bonds in sand

In this network, each silicon atom is bonded to four oxygen atoms in a **tetrahedral** structure and each oxygen atom is bonded to two silicon atoms (Fig. 7.11). The formula of silicon(IV) oxide is therefore SiO_2. It is also known as silicon dioxide.

Silicon(IV) oxide melts only at high temperatures. This is because the silicon and oxygen atoms are all held together by strong covalent bonds in the network.

Key ideas

1. The structure, properties and uses of diamond and graphite are shown below.

	Diamond	Graphite
Giant molecular structure	Each carbon atom is joined to four other carbon atoms by strong covalent bonds in a tetrahedral arrangement.	• Within each layer, each carbon atom is joined to three other carbon atoms by strong covalent bonds. • Layers of carbon atoms lie on top of each other, held together by weak forces of attraction
Properties	• Hard • Very high melting and boiling points • Non-conductor of electricity	• Soft • Very high melting and boiling points • Conductor of electricity
Uses	• As gemstones • As tips of cutting, grinding and polishing tools	• In pencils • As a dry lubricant • As inert electrodes

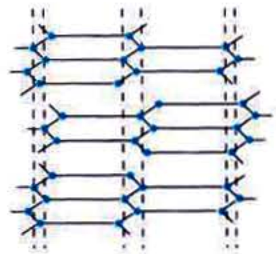
2. The table below summarises the main differences between covalent substances with simple molecular structures and those with giant molecular structures.

Type of structure	Simple molecular structure	Giant molecular structure
Examples	Hydrogen, oxygen, water, carbon dioxide, methane, iodine	Diamond, graphite, silicon (IV) oxide
Volatility	Volatile, low melting and boiling points	Non-volatile, high melting and boiling points
State (at r.t.p.)	Usually liquids or gases at room temperature	Solids at room temperature (most are hard solids, except graphite)
Solubility	Insoluble in water and usually soluble in organic solvents	Insoluble in all solvents
Electrical conductivity	Usually non-conductors of electricity in all physical states	Non-conductors of electricity (except graphite)

Test Yourself 7.2

Worked Example

What substance is represented by the structure shown in the diagram on the right?



A Diamond
B Graphite
C Silicon(IV) oxide
D Solid iodine

Thought process

Graphite has a layered structure as shown in the diagram.

Answer

B

Questions

1. Substance *X* melts at 118 °C and boils at 299 °C. Predict its
   a) solubility in water.
   b) electrical conductivity.

2. Compare ionic compounds and simple molecular covalent compounds based on these properties: melting and boiling points, solubility in water and organic solvents, and electrical conductivity.

3. Diamond and graphite are allotropes of carbon. Explain
   a) how they differ in terms of their molecular structure.
   b) why diamond is denser than graphite.

Chemical Formulae of Covalent Substances

How do we write the formulae of covalent elements?

Some covalent elements exist as diatomic molecules. To write the formula of such an element, we add a subscript '2' to the chemical symbol (Table 7.9).

For elements that are not made up of diatomic molecules, their formulae are often written as the chemical symbol of the element. For example, the formula of carbon is C and the formula of sulphur is S.

Covalent element	Chemical formula
bromine	Br_2
chlorine	Cl_2
hydrogen	H_2
iodine	I_2
nitrogen	N_2
oxygen	O_2
carbon	C
sulphur	S

Table 7.9 Chemical formulae of common covalent elements

TidBit

Unlike most covalent substances, carbon exists as a giant molecule. Sulphur actually exists as a polyatom, S_8. For simplicity, we write their formulae as C and S.

Prefix	Number of atoms
mono	1
di	2
tri	3
tetra	4
penta	5

Table 7.10 Common prefixes in naming compounds

How do we write the formulae of covalent compounds?

The name of a covalent compound will tell you what elements it contains as well as the number of each type of atom a molecule contains. Table 7.10 shows the prefixes commonly used to show the number of atoms.

Example 1

What is the chemical formula of carbon dioxide?

Solution: CO_2

carbon dioxide

Comments: We assume the first element named (in this case, carbon) contains only one atom unless otherwise stated. The word 'dioxide' means 'two oxygen atoms'.

TidBit

Dinitrogen monoxide (N_2O) is also known as 'laughing gas'. Inhaling the gas for a short time will make you numb to pain. It will even make you laugh!

Example 2

What is the chemical formula of dinitrogen monoxide?

Solution: N_2O

dinitrogen monoxide

Comments: Notice the subscript '1' is not indicated on the formula.

Example 3

What is the chemical formula of dinitrogen tetroxide?

Solution: N_2O_4

dinitrogen tetroxide

In addition to knowing these rules, you should be familiar with the chemical formulae of some common covalent substances, such as ammonia (NH_3), hydrogen peroxide (H_2O_2), methane (CH_4), ozone (O_3) and water (H_2O).

Cleaning contact lenses with hydrogen peroxide

7.3 | Metallic Bond

In ionic compounds, the ions are held together by strong ionic bonds in a giant ionic lattice. In giant covalent molecules, the atoms are held together by strong covalent bonds in a giant covalent lattice.

Metals are different. Metal atoms are held strongly to each other by **metallic bonding**. In the metal lattice, the atoms lose their valence electrons and become positively charged. The valence electrons no longer belong to any metal atom and are said to be **delocalised**. They move freely between the metal ions like a cloud of negative charge. Hence, this lattice structure is described as a *lattice of positive ions surrounded by a 'sea of mobile electrons'* (Fig. 7.12). We can therefore define a **metallic bond** as *the force of attraction between positive metal ions and the 'sea of delocalised electrons'*.

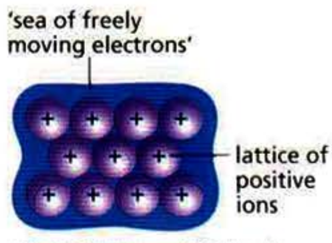


Fig. 7.12 The metallic bond

Physical Properties of Metals

Electrical conductivity

Metals are good conductors of electricity because of the mobility of the valence electrons within the metal lattice. When a metal is used in an electrical circuit, electrons entering one end of the metal cause a similar number of electrons to be displaced from the other end (Fig. 7.13). The valence electrons move from the negative terminal to the positive terminal of the electrical circuit. Hence, the metal is able to conduct an electric current.

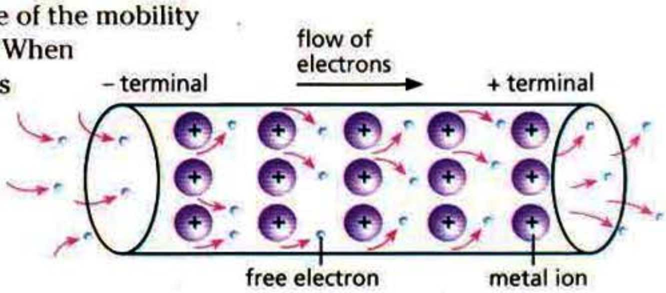
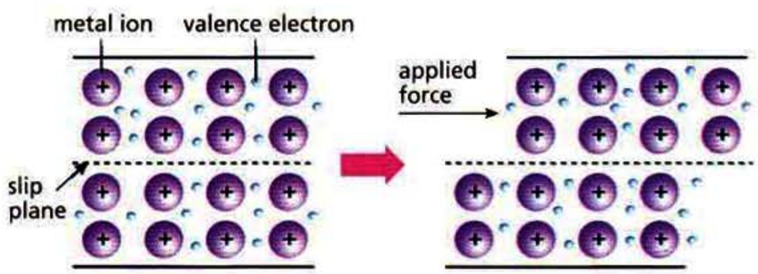


Fig. 7.13 How metals conduct electricity

Malleability and ductility

In metallic bonding, the valence electrons do not belong to any particular metal atom. If sufficient force is applied to the metal, one layer of atoms can slide over another without disrupting the metallic bonding (Fig. 7.14). As a result, metallic bonds are strong but flexible, so metals can be hammered into different shapes (malleable) or drawn into wires (ductile) without breaking.



(a) Before a force is applied, the metal atoms form an orderly and rigid arrangement.

(b) When a force is applied, one layer of metal ions slides over the other.

Fig. 7.14 Effect of force being applied on a metal lattice

Deducing the Structure and Bonding of a Substance

We can deduce the structure of a substance from its physical properties as shown in Fig. 7.15.

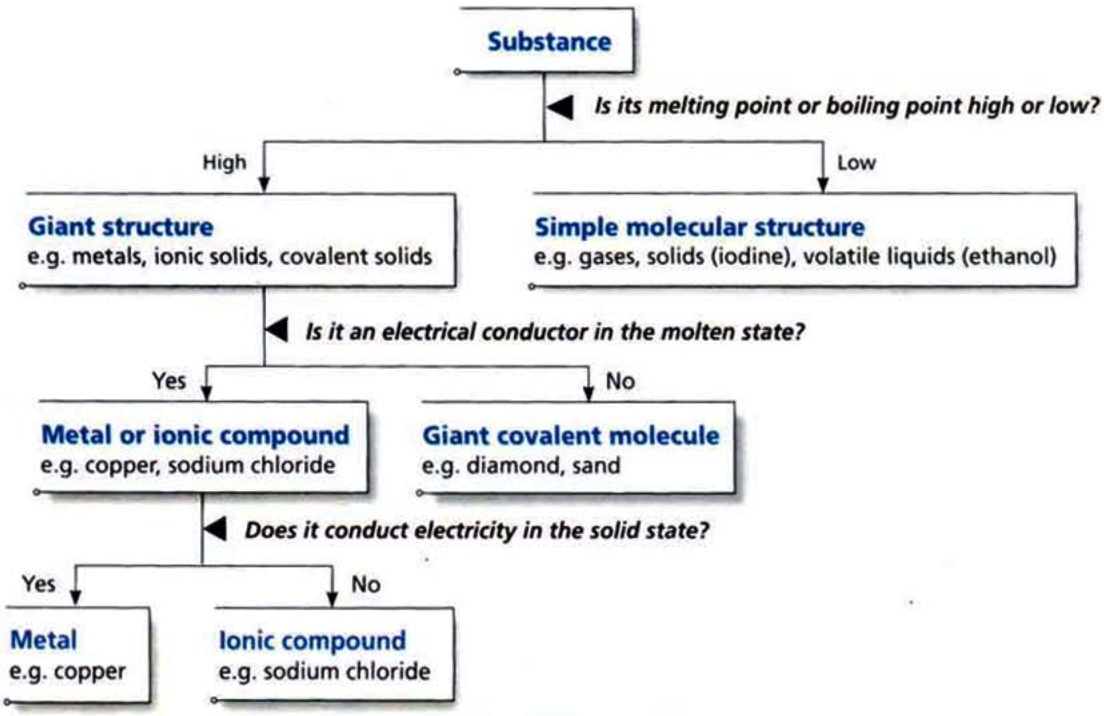


Fig. 7.15 Deducing the structure of a substance

Key ideas

Comparing different giant structures:

Type of giant structure	Structure	Typical properties				
		Melting and boiling points	Hardness	Solubility in water	Solubility in organic solvents	Electrical conductivity
Giant metallic	attraction between positive metals ions and 'sea of delocalised electrons'	high	hard but malleable	insoluble	insoluble	conductor in solid and molten states
Giant covalent	covalent bonds between all atoms	high	hard but brittle	insoluble	insoluble	non-conductor (except graphite)
Giant ionic	electrostatic attraction between positive and negative ions	high	hard but brittle	usually soluble	insoluble	non-conductor in solid state but conductor when molten or in aqueous solution

Test Yourself 7.3

Worked Example

Molten ionic compounds and aqueous solutions of ionic compounds are known as electrolytes because they can conduct electricity. Which particles are responsible for conducting electricity in electrolytes and in metals?

	Electrolytes	Metals
A	electrons	electrons
B	positive ions and negative ions	electrons
C	electrons	positive ions
D	positive ions and negative ions	positive ions and electrons

Thought Process

Electrolytes conduct electricity by the movement of positive and negative ions. Metals conduct electricity because their valence electrons can move freely between the metal ions.

Answer

B

Questions

1. Name the following covalent substances:
   a) NO
   b) SO_2
   c) N_2O_4
   d) H_2O_2
   e) NH_3

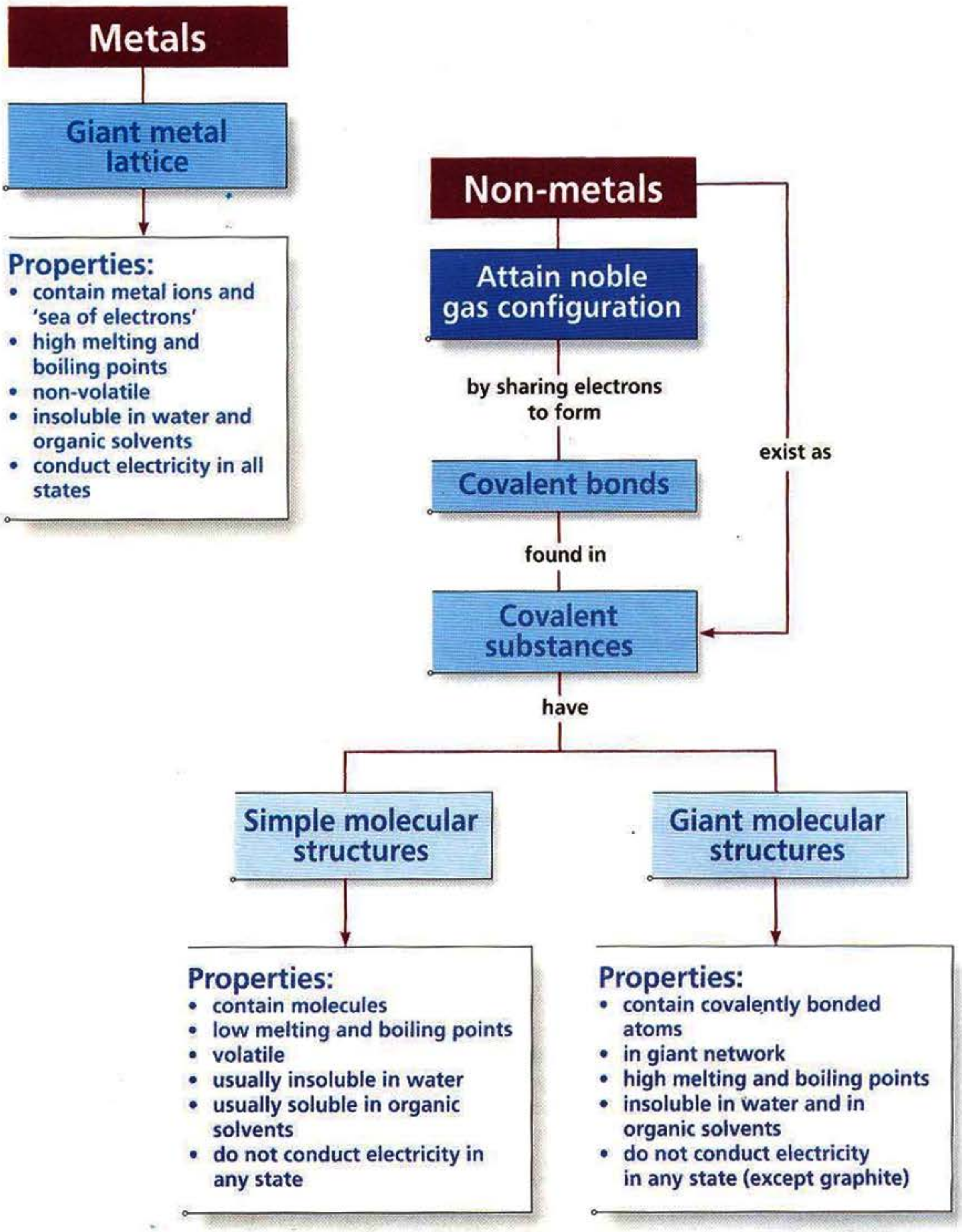
2. What is the chemical formula of each of the following substances?
   a) Nitrogen dioxide
   b) Sulphur trioxide
   c) Methane
   d) Ozone
   e) Dinitrogen monoxide

3. $NH_4Cl(s)$, $CuCl_2(s)$, $SiCl_4(l)$

   a) From the list of substances above, select one that contains both ionic and covalent bonds. Explain your choice.
   b) Predict **two** physical properties (other than melting and boiling points) of this substance. Explain your answer.

4. The melting point of element *X* is 686 °C. *X* is a good conductor of electricity in its solid and molten states. Deduce the bonding and structure of *X* and explain how they account for these properties.

Concept Map


Metals
Giant metal lattice
Properties:
• contain metal ions and 'sea of electrons'
• high melting and boiling points
• non-volatile
• insoluble in water and organic solvents
• conduct electricity in all states
Non-metals
Attain noble gas configuration
by sharing electrons to form
Covalent bonds
found in
exist as
Covalent substances
have
Simple molecular structures
Giant molecular structures
Properties:
• contain molecules
• low melting and boiling points
• volatile
• usually insoluble in water
• usually soluble in organic solvents
• do not conduct electricity in any state
Properties:
• contain covalently bonded atoms
• in giant network
• high melting and boiling points
• insoluble in water and in organic solvents
• do not conduct electricity in any state (except graphite)

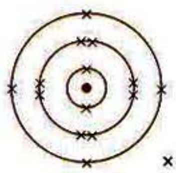
Exercise 7

Foundation

1. The table below shows the properties of sodium chloride and hydrogen chloride. Which property is wrong?

	Sodium chloride	Hydrogen chloride
A	conducts electricity when dissolved in water	conducts electricity when dissolved in water
B	covalently bonded	ionically bonded
C	solid at room temperature	gas at room temperature
D	soluble in water	soluble in water

2. a) The electronic arrangement of an atom of a non-metallic element *Y* is shown below.



×: electron of *Y*

Atom of element *Y*

i) Draw a 'dot and cross' diagram to show the electronic arrangement of an atom of chlorine. The atomic number of chlorine is 17.

ii) Draw a 'dot and cross' diagram to show the electronic arrangement of the compound formed when *Y* and chlorine react.

iii) What type of chemical bond is formed in (ii)?

b) Use 'dot and cross' diagrams to show the bonding between sulphur and chlorine in sulphur dichloride, SCl_2.

Challenge

1. H_2S and HCl are two covalent compounds. What is the most likely structural formula for the covalent compound S_2Cl_2?

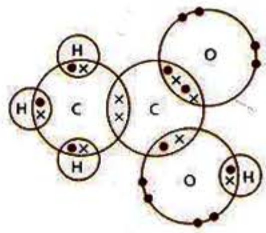
A Cl–Cl–S–S B Cl–S–S–Cl
C S–Cl–Cl–S D S–Cl–S–Cl

2. Common salt, $NaCl$, is a solid and methane, CH_4, is a gas at room temperature.

a) Draw 'dot and cross' diagrams to show the bonding present in both compounds.

b) Explain, in terms of the bonds present in each compound, why their physical states are different.

3. The diagram below shows the arrangement of electrons in an organic compound *X*.



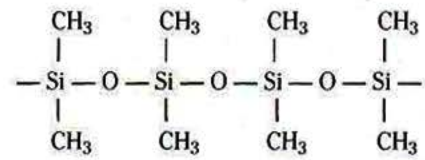
a) Name the elements present in *X*.

b) Which two atoms have a double bond between them?

c) How many single covalent bonds can each atom of C form?

d) Suggest, with reasons, **three** properties of compound *X*.

4. a) Silicone is used extensively in the chemical industry. It has the formula:

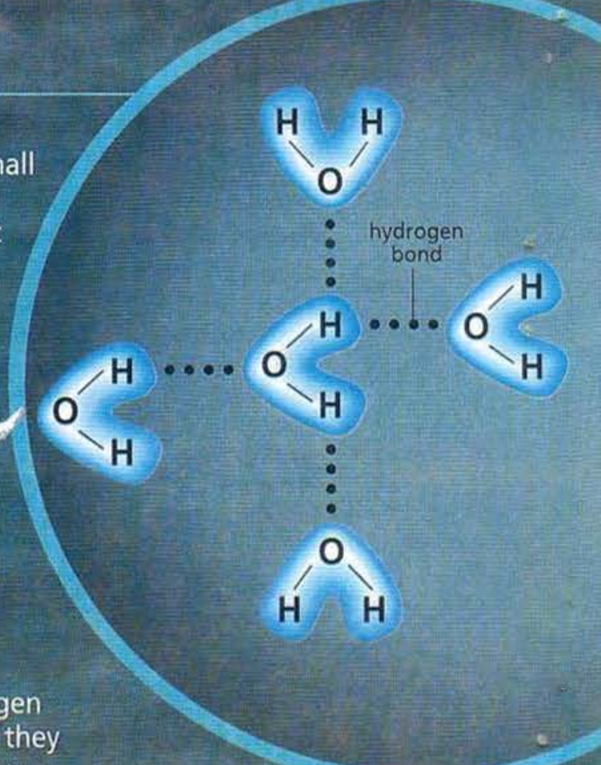


Predict **two** physical properties of silicone.

b) Carbon dioxide is a gas at room temperature and silicon(IV) oxide is a solid with a high melting point. Explain these observations in terms of the structure and bonding of both substances.

Chemistry Today

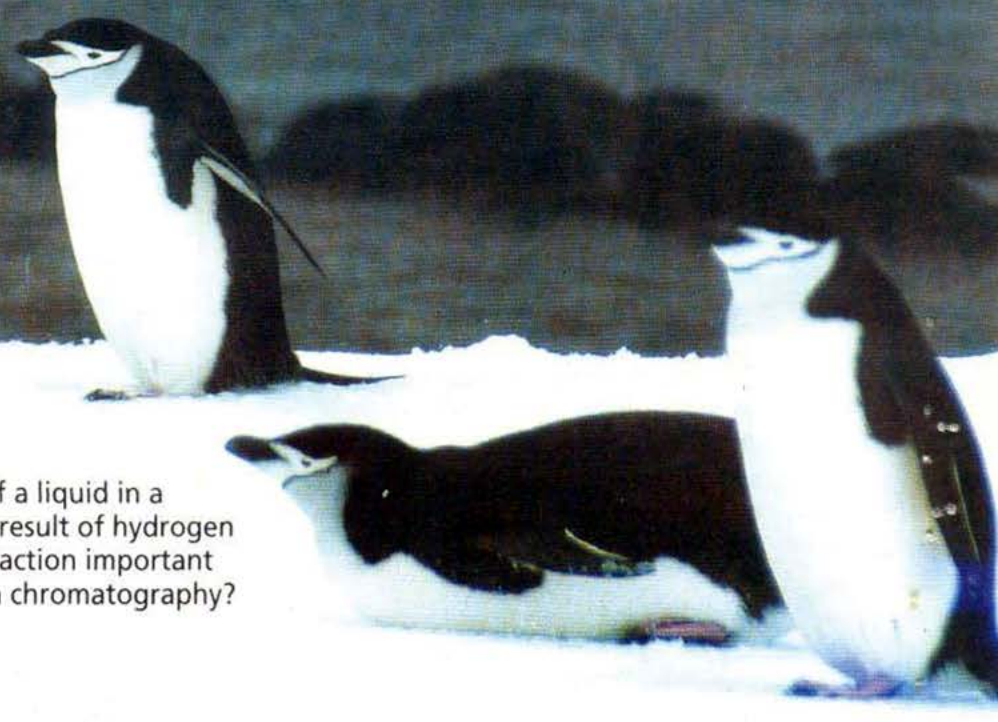
Have you ever wondered why water is a liquid when other small covalent molecules such as carbon dioxide, chlorine, oxygen and nitrogen are gases? What property does water have that causes it to be almost non-volatile at room temperature and pressure?



Water is a covalent molecule, but there is another type of bond holding the water molecules together called a hydrogen bond, represented by a dotted line in the figure shown. The hydrogen atoms on one water molecule are attracted to the oxygen atom of another water molecule to form a much larger molecule. (The diagram shows that one water molecule is bonded to four other water molecules, but in fact thousands of such molecules are held together by hydrogen bonding!)

Hydrogen bonding is very noticeable in ice. Because of hydrogen bonding, the molecules of water in ice are further apart than they are in liquid water. This causes ice to have a lower density than water, hence ice floats on water.

Ice is less dense than water, so it floats over lakes, providing insulation and thus preventing lakes from freezing entirely. Thus aquatic creatures can survive freezing temperatures.

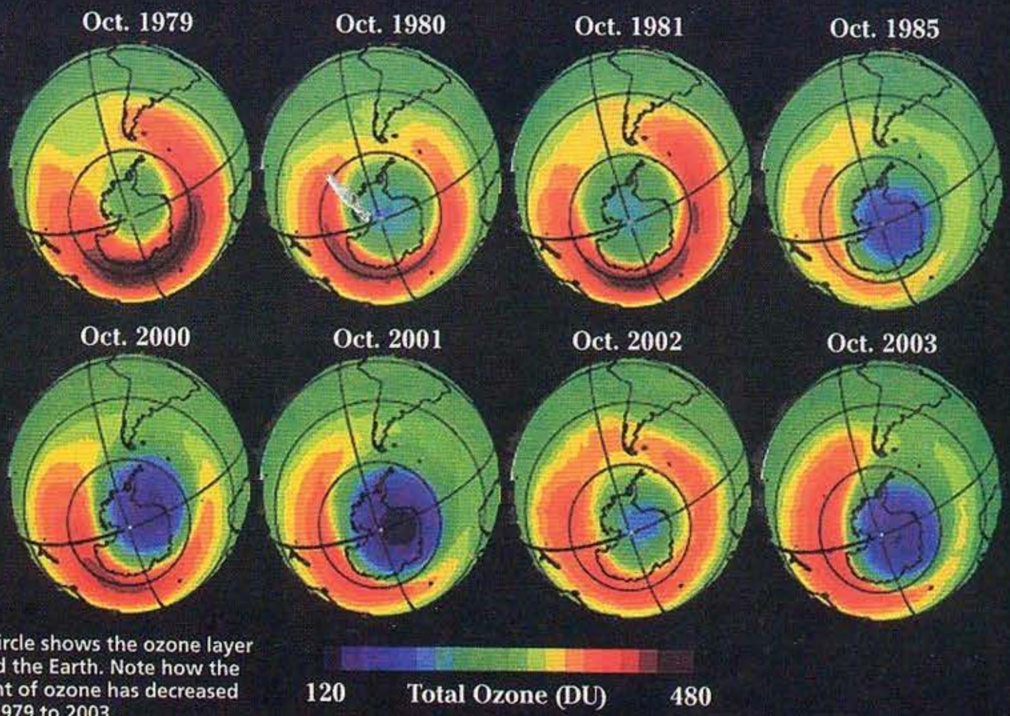


CRITICAL THINKING

Capillary attraction (the rise of a liquid in a tube of small diameter) is the result of hydrogen bonding. Why is capillary attraction important (i) in the growth of trees (ii) in chromatography?

Chapter **8**

Writing Equations



Each circle shows the ozone layer around the Earth. Note how the amount of ozone has decreased from 1979 to 2003.

A chemical called ozone exists in our stratosphere. It forms a protective layer around the Earth. It absorbs harmful radiation from the Sun, reducing our risk of getting skin cancer.

Unfortunately, the ozone layer around the Earth is disappearing. Scientists performed experiments and discovered that some chemicals that we release into the atmosphere react with ozone and break it down. These reactions, as well as other chemical reactions, can be represented as chemical equations.

Chemical equations can be used to represent how the ozone layer is disappearing. They are also important in many other situations, as you will see in the next two chapters. In this chapter, you will learn how to write chemical equations correctly.

Chapter Outline

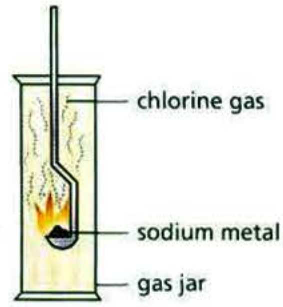


Fig. 8.1 Sodium and chlorine react to form sodium chloride.

8.1 | Chemical Equations

What is a chemical equation?

In chemistry, an equation represents what occurs in a chemical reaction. We can write an equation by using words or chemical formulae. The most important thing to remember is that an equation must represent a reaction known to take place. The **word equation** (equation in words) for the reaction between sodium and chlorine is written below.

Word equation

sodium + chlorine ⟶ sodium chloride

We can also *make use of chemical formulae to write* the **chemical equation.**

Chemical equation

$Na + Cl_2 \longrightarrow NaCl$

We write the reacting substances or **reactants** on the left-hand side of the equation.

In this case, sodium (Na) and chlorine (Cl_2) are the reactants.

The arrow means 'react to form'.

The reaction proceeds from left to right.

We write the substances formed or **products** on the right-hand side of the equation.
Here, sodium chloride (NaCl) is the only product.

Notice that on the left-hand side of the equation, there are two chlorine atoms but on the right-hand side of the equation, there is only one chlorine atom. This means that the equation is *not* balanced. An equation is balanced when there is an *equal number of atoms of each element on both sides of the equation*. In the next section, you will learn how to write a **balanced chemical equation**.

Chem-Aid

(s) represents a solid.
(*l*) represents a liquid.
(g) represents a gas.
(aq) represents an aqueous solution, i.e. dissolved in water.

Writing Balanced Equations

An equation is usually written using chemical formulae, unless you are instructed to write a word equation. There are four steps in writing a balanced chemical equation:

Step	
1	Write down the formulae of the reactants and products.
2	Check the number of atoms of each element on both sides of the equation. If the equation is not balanced, proceed to step 3.
3	Balance the equation by placing numbers in front of the formulae of the substances in the equation. The number '1' is *not* written.
4	Include the state symbols in the equation.

Did you know that common salt is actually sodium chloride?

Example 1

Making use of these four steps, we can write the balanced chemical equation for the reaction between sodium and chlorine.

Solution:

Step		
1	Write down the formulae of the reactants and products.	$Na + Cl_2 \longrightarrow NaCl$ sodium chlorine sodium chloride
2	Check the number of atoms of each element on both sides of the equation.	There are two chlorine atoms on the left-hand side but only one chlorine atom on the right-hand side. This means that the equation is not balanced.
3	Balance the equation. To balance the number of chlorine atoms, we need to put a '2' in front of the 'NaCl'. The equation is still not balanced because there are two sodium atoms on the right-hand side and one sodium atom on the left-hand side. To balance the equation, we need to put a '2' in front of 'Na'.	$Na + Cl_2 \longrightarrow 2NaCl$ $2Na + Cl_2 \longrightarrow 2NaCl$
4	Add the state symbols.	$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

The balanced equation reads as 'Two atoms of sodium react with one molecule of chlorine to form two formula units of sodium chloride'.

Example 2

Methane gas (CH_4) is often used as fuel for Bunsen burners in laboratories. Methane burns in oxygen to form carbon dioxide and water vapour. Write an equation for this reaction.

Chem-Aid

When balancing equations, never change the formula of a substance.

Solution:

Step			
1	Write down the formulae of the reactants and products.	$CH_4 + O_2 \longrightarrow CO_2 + H_2O$ methane oxygen carbon dioxide water vapour	
2	Check the number of atoms of each element on both sides of the equation.	**reactants**	**products**
		1 carbon atom	1 carbon atom
		4 hydrogen atoms	2 hydrogen atoms
		2 oxygen atoms	3 oxygen atoms
3	Balance the equation.	$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$	
4	Add the state symbols. Note that the state symbol for H_2O is 'g' because water vapour (a gas) is formed.	$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$	

What does the balanced equation read as?

Test Yourself 8.1

Question

Copper(II) oxide reacts with ammonia to give copper, water and nitrogen. What must the values of *x*, *y* and *z* be in order to balance this equation?

$$3CuO(s) + 2NH_3(g) \longrightarrow xCu(s) + yH_2O(l) + zN_2(g)$$

8.2 | Ionic Equations

Many substances, especially ionic compounds, are soluble in water. An **ionic equation** is a simplified chemical equation that shows the reactions of such substances in water.

How do we write ionic equations?

When aqueous sodium chloride is added to aqueous silver nitrate, a white precipitate of silver chloride is formed. This reaction can be represented by the chemical equation below.

Word equation sodium chloride + silver nitrate ⟶ silver chloride + sodium nitrate

Chemical equation $NaCl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$

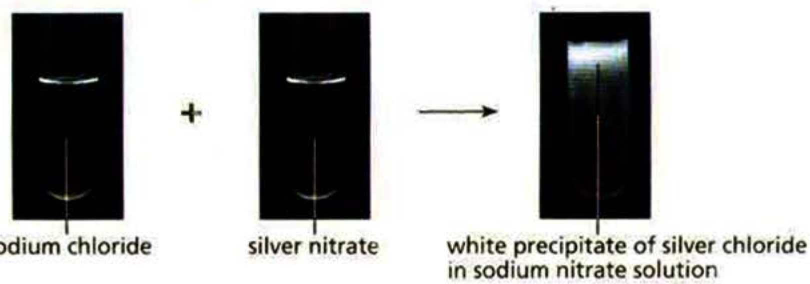


Fig. 8.2 This reaction produced silver chloride precipitate.

Sodium chloride, silver nitrate and sodium nitrate are soluble in water. They exist as ions in solution. If we write the chemical equation in terms of ions, we get

$$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \longrightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

The ions marked in red (Na^+ ions and NO_3^- ions) have *not* taken part in any chemical reactions. They are still ions in solution at the end of the reaction. Such ions are called **spectator ions**.

Since only silver ions and chloride ions have reacted, the equation for the reaction can therefore be simplified as shown below.

Ionic equation $Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

The above equation is called an **ionic equation**. An **ionic equation** *shows the ions taking part in a reaction and the products formed. It leaves out the spectator ions that do not react.*

How do we write an ionic equation?
The steps in writing an ionic equation are:

Step	
1	Write the balanced chemical equation of the reaction. Include the state symbols.
2	Identify ionic compounds that are soluble in water (you may refer to page 196 to check solubility of compounds). These compounds become ions in water. Rewrite the chemical equation in terms of ions.
3	Cancel out the spectator ions.
4	Write the ionic equation.

Chem-Aid

Some covalent compounds form ions in water, e.g. hydrochloric acid (HCl).

Example 1
When hydrochloric acid reacts with sodium hydroxide solution, the products formed are sodium chloride and water. Write the ionic equation for this reaction.

Solution:

Step		
1	Write the balanced chemical equation.	$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$
2	Rewrite the equation in terms of ions for substances that are soluble in water.	$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$
3	Cancel out the spectator ions.	$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \longrightarrow \cancel{Na^+(aq)} + \cancel{Cl^-(aq)} + H_2O(l)$
4	Write the ionic equation.	$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

Link

Sodium hydroxide ($NaOH$) is an alkali. Find out more about alkalis and acids in chapter 11.

Example 2
When zinc is added to copper(II) sulphate solution, copper and zinc sulphate are formed. Write the ionic equation for this reaction.

Solution:

Step		
1	Write the balanced chemical equation.	$Zn(s) + CuSO_4(aq) \longrightarrow Cu(s) + ZnSO_4(aq)$
2	Rewrite the equation in terms of ions for substances that are soluble in water.	$Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow Cu(s) + Zn^{2+}(aq) + SO_4^{2-}(aq)$
3	Cancel out the spectator ions.	$Zn(s) + Cu^{2+}(aq) + \cancel{SO_4^{2-}(aq)} \longrightarrow Cu(s) + Zn^{2+}(aq) + \cancel{SO_4^{2-}(aq)}$
4	Write the ionic equation.	$Zn(s) + Cu^{2+}(aq) \longrightarrow Cu(s) + Zn^{2+}(aq)$

Example 3
A solution of sodium carbonate reacts with sulphuric acid to form sodium sulphate, carbon dioxide and water. Write the ionic equation for this reaction.

Solution:

Step		
1	Write the balanced chemical equation.	$Na_2CO_3(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(l)$
2	Rewrite the equation in terms of ions for substances that are soluble in water.	$2Na^+(aq) + CO_3^{2-}(aq) + 2H^+(aq) + SO_4^{2-}(aq) \longrightarrow 2Na^+(aq) + SO_4^{2-}(aq) + CO_2(g) + H_2O(l)$
3	Cancel out the spectator ions.	$\cancel{2Na^+(aq)} + CO_3^{2-}(aq) + 2H^+(aq) + \cancel{SO_4^{2-}(aq)} \longrightarrow \cancel{2Na^+(aq)} + \cancel{SO_4^{2-}(aq)} + CO_2(g) + H_2O(l)$
4	Write the ionic equation.	$CO_3^{2-}(aq) + 2H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$

Key ideas

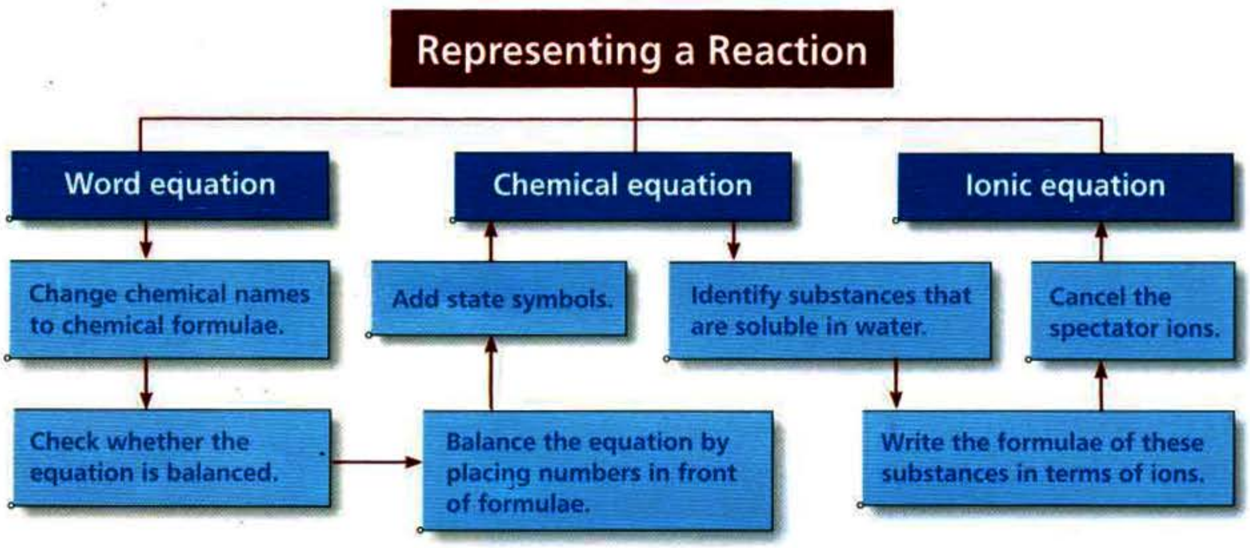
1. A chemical equation is used to represent a chemical reaction.
2. A chemical equation is balanced when there is an equal number of atoms of each element on both sides of the equation.
3. An ionic equation shows the reaction between ions.

Test Yourself 8.2

Questions

1. Which of the following equations are balanced?
   a) $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
   b) $2Al(s) + Fe_2O_3(s) \longrightarrow 2Fe(s) + Al_2O_3(s)$
   c) $4NO_2(g) + H_2O(l) + 2O_2(g) \longrightarrow 4HNO_3(l)$
2. Identify the spectator ions in the following reactions.
   a) $Na_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow BaSO_4(s) + 2NaNO_3(aq)$
   b) $CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Concept Map



Exercise 8

Foundation

1. Phosphorus(V) chloride dissolves in water to form phosphoric acid, H_3PO_4 and hydrochloric acid. What is the value of x, y and z in the balanced equation for this reaction?

$$PCl_5(s) + xH_2O(l) \longrightarrow yH_3PO_4(aq) + zHCl(aq)$$

	x	y	z
A	2	1	4
B	2	2	5
C	4	2	4
D	4	1	5

2. Sodium reacts with water to produce sodium hydroxide and hydrogen gas. Which of the following shows the correct chemical equation?

A $Na(s) + H_2O(l) \longrightarrow NaOH(aq) + \frac{1}{2}H_2(g)$
B $Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$
C $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$
D $2Na(s) + H_2O(l) \longrightarrow Na_2OH(aq) + H_2(g)$

3. Write balanced chemical equations for each of the following reactions. State symbols are not required.
   a) calcium + chlorine $\longrightarrow$ calcium chloride
   b) hydrogen + nitrogen $\longrightarrow$ ammonia
   c) ethane (C_2H_6) + oxygen $\longrightarrow$ carbon dioxide + water vapour

4. Balance the following equations and include the state symbols. You may check the solubility of the substances on page 196.
   a) $Fe + Cl_2 \longrightarrow FeCl_3$
   b) $H_2SO_4 + Al(OH)_3 \longrightarrow Al_2(SO_4)_3 + H_2O$
   c) $Pb^{2+}(aq) + Cl^- \longrightarrow PbCl_2$
   d) $H^+(aq) + CO_3^{2-} \longrightarrow H_2O + CO_2$

Challenge

1. Potassium hydroxide (KOH) and sulphuric acid (H_2SO_4) react to produce potassium sulphate (K_2SO_4) and water. Which of the following shows the ionic equation for the reaction?

A $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$
B $H^+(aq) + OH^-(aq) \longrightarrow \frac{1}{2}H_2O(l)$
C $2K^+(aq) + SO_4^{2-}(aq) \longrightarrow K_2SO_4(s)$
D $K^+(aq) + SO_4^{2-}(aq) + H^+(aq) \longrightarrow KHSO_4(s)$

2. When aqueous sodium phosphate, Na_3PO_4, was added to aqueous calcium chloride, $CaCl_2$, calcium phosphate was precipitated.
   a) Write the formula for the ionic compound, calcium phosphate, formed.
   b) i) Write the chemical equation for the reaction between calcium chloride and sodium phosphate.
   ii) Write the ionic equation for the reaction in (b)(i).

Chemistry Today

At the start of this chapter, we learnt that the ozone layer is disappearing. The main cause of the loss of ozone seems to be the reaction between chlorofluorocarbons (CFCs) and ozone. CFCs were found in aerosol cans. CFCs were released when aerosols were used or when air conditioners and freezers were not disposed of properly. Most countries have banned the use of CFCs, but the amount of CFCs already in the atmosphere will last for a long time.

CFCs rise into the atmosphere and release chlorine atoms. These chlorine atoms react with ozone (O_3) and produce oxygen. What is happening can be represented by a word equation:

ozone + chlorine atom
⟶ oxygen + compound of chlorine and oxygen

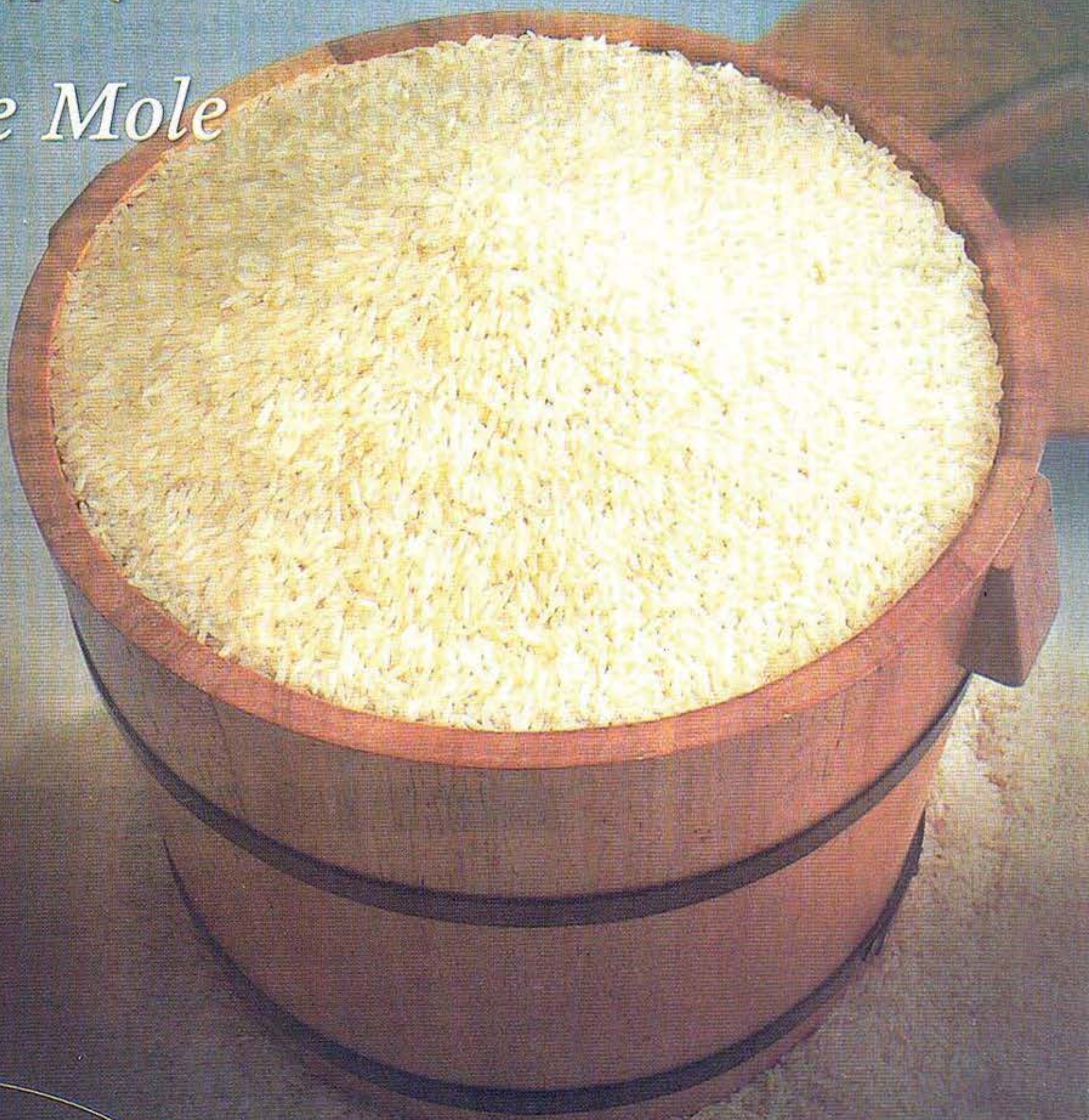
As a result, the amount of ozone to protect us from harmful radiation is decreasing.

CRITICAL THINKING

Another chemical that is destroying the ozone layer is nitrogen monoxide (NO) which reacts with ozone in the same way as CFCs. Write a word equation to show the reaction between nitrogen monoxide and ozone.

Chapter 9

The Mole



Chapter Outline

Have you ever tried counting the number of rice particles in a bucket of rice? It is difficult to do so because rice particles are very small and numerous. Chemists face a similar problem when they try to count atoms. Atoms are too small to be counted one at a time. Because they are so small, it is also difficult to measure the mass of each atom. In this chapter, you will find out how chemists overcome the problems of counting atoms and measuring the mass of each atom.

TidBit

The mass of a single carbon atom is about 0.000 000 000 000 000 000 000 000 020 04 g or 2.004×10^{-29} kg!

Chem-Aid

The carbon-12 atom has six protons, six neutrons and six electrons.

Link

Recall what you have learnt about isotopes in chapter 5.

9.1 | Relative Atomic Mass

How can we measure the mass of an atom?

You learnt in chapter 4 that atoms are very small particles. Atoms also have very small masses, so it is not practical to use the actual masses of atoms in calculations.

To overcome this problem, chemists often compare masses of different atoms with the carbon-12 atom (an isotope of carbon). Scientists all over the world agreed to give the carbon-12 atom a relative atomic mass of 12. The masses of all other atoms are compared with one-twelfth the mass of one carbon-12 atom.

The **relative atomic mass** of any atom is *the number of times the mass of one atom of an element is greater than $\frac{1}{12}$ of the mass of one carbon-12 atom.*

$$\text{Relative atomic mass} = \frac{\text{mass of one atom of the element}}{\text{mass of } \frac{1}{12} \text{ of an atom of carbon-12}}$$

For example, one atom of oxygen is 16 times heavier than $\frac{1}{12}$ of an atom of carbon-12. Thus, oxygen has a relative atomic mass of 16.

The symbol for relative atomic mass is $\mathbf{A_r}$. Relative atomic mass is a ratio and therefore has *no unit*. The relative atomic masses of elements are given in the Periodic Table.

Element	Relative atomic mass, A_r
hydrogen	1
carbon	12
oxygen	16
chlorine	35.5

Table 9.1 Relative atomic masses of some common elements

Why are some A_r values not whole numbers?

As you can see from Table 9.1, the relative atomic mass of an element is usually a whole number. However, the relative atomic masses of some elements, such as chlorine, are not whole numbers. This is because such elements occur as mixtures of isotopes.

For example, chlorine exists in two isotopic forms: chlorine-35 and chlorine-37. A sample of chlorine is made up of 75% of chlorine-35 atoms and 25% of chlorine-37 atoms.

Hence, relative atomic mass of chlorine

$$= (\frac{75}{100} \times 35) + (\frac{25}{100} \times 37)$$
$$= 26.25 + 9.25$$
$$= 35.5$$

9.2 | Relative Molecular Mass

Many elements and compounds exist as molecules. For example, chlorine exists as molecules. Each molecule of chlorine consists of two chlorine atoms. One molecule of nitrogen dioxide consists of one nitrogen atom and two oxygen atoms (Fig. 9.1).

One molecule of chlorine   One molecule of nitrogen dioxide

Fig. 9.1 Chlorine and nitrogen dioxide exist as molecules.

The mass of a molecule is measured in terms of its relative molecular mass. The **relative molecular mass (M_r)** of an element or compound is *the mass of a molecule*, compared to $\frac{1}{12}$ *the mass of one atom of carbon-12.*

Relative molecular mass (M_r)

$$= \frac{\text{mass of one molecule of an element or compound}}{\text{mass of } \frac{1}{12} \text{ of an atom of carbon-12}}$$

How do we calculate the relative molecular mass of a molecule?

The relative molecular mass of a molecule is calculated by adding together the relative atomic masses of each atom in its chemical formula (Table 9.2). Like relative atomic mass, it is a ratio and therefore has *no unit.*

Molecule	Chemical formula	Number of atoms in one molecule	Calculating M_r
nitrogen	N_2	2 N	$(2 \times 14) = \mathbf{28}$
ammonia	NH_3	1 N; 3 H	$(1 \times 14) + (3 \times 1) = \mathbf{17}$
carbon dioxide	CO_2	1 C; 2 O	$(1 \times 12) + (2 \times 16) = \mathbf{44}$
water	H_2O	2 H; 1 O	$(2 \times 1) + (1 \times 16) = \mathbf{18}$
ethanol	C_2H_5OH	2 C; 6 H; 1 O	$(2 \times 12) + (6 \times 1) + (1 \times 16) = \mathbf{46}$

Table 9.2 Calculating the relative molecular masses of some molecules

Relative Formula Mass

You learnt in chapter 7 that substances like water are covalent and exist as molecules. However, substances like sodium chloride are ionic and do not exist as molecules.

The relative molecular mass of an ionic compound is more accurately known as the **relative formula mass**. Like relative molecular mass, relative formula mass is given the symbol M_r and has *no units*. For example, the relative formula mass of sodium chloride (NaC*l*) is 23 + 35.5 = 58.5.

Substance	Formula unit	Number of atoms in formula unit	Calculating M_r
magnesium sulphate	$MgSO_4$	1 Mg; 1 S; 4 O	$(1 \times 24) + (1 \times 32) + (4 \times 16) = \mathbf{120}$
calcium carbonate	$CaCO_3$	1 Ca; 1 C; 3 O	$(1 \times 40) + (1 \times 12) + (3 \times 16) = \mathbf{100}$
calcium nitrate	$Ca(NO_3)_2$	1 Ca; 2 N; 6 O	$(1 \times 40) + (2 \times 14) + (6 \times 16) = \mathbf{164}$
copper(II) sulphate crystals	$CuSO_4.5H_2O$	1 Cu; 1 S; 9 O; 10 H	$(1 \times 64) + (1 \times 32) + (9 \times 16) + (10 \times 1) = \mathbf{250}$

Table 9.3 Calculating the relative formula masses of some ionic substances

Chem-Aid

The dot ('.') in $CuSO_4.5H_2O$ means that there are five H_2O molecules bonded to each $CuSO_4$. It does not mean multiply.

Key ideas

1. The relative atomic mass (A_r) of any atom is the number of times the mass of one atom of an element is heavier than $\frac{1}{12}$ of a carbon-12 atom.
2. A_r values may not be whole numbers because of isotopes.
3. The relative molecular mass (M_r) of a molecule is the sum of the relative atomic masses of all the atoms in the molecule.
4. The relative formula mass (M_r) of an ionic compound is the sum of the relative atomic masses of atoms in a formula unit of the compound.

Test Yourself 2.1

Worked Example

$^{24}_{12}S$ and $^{16}_{8}T$ are two elements that react together to form an ionic compound *V*. What is the relative formula mass (M_r) of *V*?

A 20 B 28 C 32 D 40

Thought Process

The electronic configuration of *S* is (2, 8, 2). *S* has a valency of 2. The electronic configuration of *T* is (2, 6). *T* also has a valency of 2. Hence, the formula of this compound is *ST*. Accordingly, the M_r of *V* is 24 + 16 = 40.

Answer

D

Questions

1. Calculate the relative molecular mass/relative formula mass of each of the following.
   a) $Ca(OH)_2$
   b) $(NH_4)_3PO_4$
   c) $C_2H_5COOCH_3$
   d) $Pb(NO_3)_2$
   e) $Na_2CO_3.10H_2O$
2. Four nitrogen atoms have the same mass as one formula unit of *XO*. What is *X*?

9.3 | The Mole

Atoms are too small and numerous to be counted one at a time. Instead, the quantity of atoms is measured by mass. The unit of measurement for atoms and molecules is the **mole**. The mole is also the S.I. unit for chemical quantity. The symbol for the mole is **mol**.

A mole of substance contains the *same number of particles as the number of atoms in 12 g of carbon-12.*

How many particles are there in a mole?

There are *approximately* 6×10^{23} particles in one mole of substance. 6×10^{23} is called the **Avogadro's constant**. One mole of particles contains 6×10^{23} particles. The particles could be atoms, molecules, ions or electrons.

Chem-Aid

6×10^{23} is known as Avogadro's constant. It may also be known as Avogadro's number.

One mole of particles contains 6×10^{23} particles.

Even if you started counting now at the rate of 10 million atoms per second, it would take you 2 billion years to count up to Avogadro's number!

How do we convert between number of moles and number of particles?

Since one mole of substance contains 6×10^{23} particles,

$$\text{Number of moles} = \frac{\text{number of particles}}{6 \times 10^{23}}$$

Equal numbers of moles contain equal numbers of particles. The reverse is also true.

Example 1

Convert 1×10^{23} of neon atoms to moles of neon atoms.

Solution:

$$\begin{aligned}\text{Number of moles of neon atoms} &= \frac{\text{number of neon atoms}}{\text{Avogadro's constant}}\\ &= \frac{1 \times 10^{23}}{6 \times 10^{23}}\\ &= \mathbf{0.167\ mol}\end{aligned}$$

Example 2

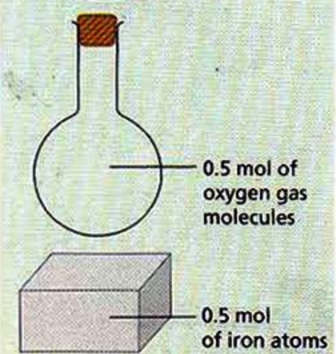
How many iron atoms are there in 0.5 mol of iron?

Solution:

In one mole of iron, there are 6×10^{23} iron atoms.

$$\begin{aligned}\text{Number of iron atoms} &= \text{number of moles} \times 6 \times 10^{23}\\ &= 0.5 \times 6 \times 10^{23}\\ &= \mathbf{3 \times 10^{23}}\end{aligned}$$

Quick Check



Are the number of particles in 0.5 mol of oxygen gas molecules and 0.5 mol of iron atoms the same?

Example 3

How many hydrogen atoms are there in three moles of hydrogen gas?

Solution:

Hydrogen gas is made up of hydrogen molecules (H_2).
In one mole of H_2 molecules, there are two moles of H atoms.
In three moles of H_2 molecules, there are six moles of H atoms.

Therefore, the number of hydrogen atoms
$= 6 \times 6 \times 10^{23}$
$= \mathbf{3.6 \times 10^{24}}$

How is the mole related to chemical equations?

Consider the following chemical equation:

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

It tells us that: two particles (atoms) of sodium → react with → one particle (molecule) of chlorine → to form → two units of sodium chloride

OR

in terms of moles: two moles of sodium atoms → react with → one mole of chlorine molecules → to form → two moles of sodium chloride

Understanding that equal numbers of moles contain equal numbers of particles is important when we perform chemical calculations. You will learn more about this in the next chapter.

9.4 | Mole and Molar Mass

What is the mass of one mole of atoms of an element?

The **molar mass** of an element is the *mass of one mole* of atoms of the element. Look at Table 9.4. Do you notice a relationship between the value of A_r and the molar mass of a substance?

Element	A_r	Molar mass
aluminium	27	27 g
carbon	12	12 g
neon	20	20 g
oxygen	16	16 g

Table 9.4 The molar masses of some elements

The molar mass is equal to the relative atomic mass (A_r) of the element in grams. For example, the relative atomic mass of sodium is 23. The mass of one mole of sodium atoms is 23 g. We can also say that the molar mass of sodium is 23 g.

Chem-Aid

The unit for molar mass may be 'g' or 'g/mol'.

What is the relationship between mole and molar mass?

The number of moles of an element can be calculated using the formula:

$$\text{Number of moles of an element} = \frac{\text{mass of element in grams}}{\text{relative atomic mass of the element}}$$

Example 1

Determine the number of moles in 0.196 kg of iron. (A_r: Fe = 56)

Solution:

$$\text{Number of moles of iron} = \frac{\text{mass of iron in grams}}{A_r \text{ of iron}}$$

$$= \frac{0.196 \times 1000}{56}$$

$$= \mathbf{3.5\ mol}$$

Chem-Aid

Convert the unit of mass to grams. In this example, the mass is given in kg, therefore it has to be multiplied by 1000 to convert it to grams.

Example 2

a) How many moles of lead are there in 1.204×10^{22} atoms of lead?
b) What is the mass of 1.204×10^{22} atoms of lead? (A_r: Pb = 207)

Solution:

a) $$\text{Number of moles of lead} = \frac{\text{number of atoms of lead}}{\text{Avogadro's constant}}$$

$$= \frac{1.204 \times 10^{22}}{6 \times 10^{23}}$$

$$= \mathbf{0.02\ mol}$$

b) $$\text{Mass of lead} = \text{number of moles} \times A_r \text{ of lead}$$

$$= 0.02 \times 207$$

$$= \mathbf{4.14\ g}$$

Quick Check

Which has the greater mass, 1.0 g of gold or 0.5 mol of helium? Show your working.

What is the mass of one mole of molecules or one mole of a compound?

In the previous section, you learnt that the mass of one mole of atoms is the same as its A_r in grams. The same idea can be extended to molecules and compounds. Table 9.5 shows the molar masses of some common elements and compounds. One mole of a substance will have a mass equal to the relative molecular mass or relative formula mass in grams.

Substance	Formula	M_r	Molar mass
oxygen	O_2	$2 \times 16 = 32$	32 g
iodine	I_2	$2 \times 127 = 160$	254 g
magnesium oxide	MgO	$24 + 16 = 40$	40 g
water	H_2O	$(2 \times 1) + 16 = 18$	18 g

Table 9.5 The molar masses of common substances

One mole of iodine (I_2) weighs 254 g.

The number of moles of a substance can be calculated using the formula:

$$\text{Number of moles} = \frac{\text{mass of substance (g)}}{\text{relative molecular mass}}$$

OR

$$= \frac{\text{mass of substance (g)}}{\text{relative formula mass}}$$

Remember, we have also learnt that

$$\text{Number of moles} = \frac{\text{number of particles}}{\text{Avogadro's constant}}$$

Example 1

A conical flask contains 68.4 g of octane (C_8H_{18}). How many molecules of octane are there in the flask?

Solution:

$$\text{Number of moles of octane} = \frac{\text{mass of octane}}{M_r \text{ of octane}}$$

$$= \frac{68.4}{114}$$

$$= \mathbf{0.6\ mol}$$

Number of molecules of octane

= number of moles × Avogadro's constant

$= 0.6 \times 6 \times 10^{23}$

$= \mathbf{3.6 \times 10^{23}}$

Example 2

How many **ions** are there in 20 g of magnesium oxide (MgO)?

Solution:

Number of moles of magnesium oxide

$$= \frac{\text{mass of magnesium oxide}}{M_r \text{ of magnesium oxide}}$$

$$= \frac{20}{40}$$

$$= \mathbf{0.5\ mol}$$

$$MgO \longrightarrow Mg^{2+} + O^{2-}$$

From the equation, we can see that 1 mol of MgO contains 1 mol of Mg^{2+} ions and 1 mol of O^{2-} ions. 0.5 mol of MgO will contain 0.5 mol of Mg^{2+} ions and 0.5 mol of O^{2-} ions.

Hence, 0.5 mol of MgO contains 1 mol of ions.

Number of ions $= 1 \times 6 \times 10^{23} = \mathbf{6 \times 10^{23}}$

Key ideas

1. The mass of one mole of a substance is its
   a) relative atomic mass in grams if the substance exists as atoms e.g. neon (Ne).
   b) relative molecular mass in grams if the substance exists as molecules e.g. oxygen (O_2).
   c) relative formula mass in grams if the substance is an ionic compound e.g. magnesium oxide (MgO).
2. A mole of any substance contains 6×10^{23} particles. This number is called the Avogadro's constant or Avogadro's number.
3. $$\text{Number of moles of atoms} = \frac{\text{mass of element (g)}}{A_r}$$
   $$\text{Number of moles of substance} = \frac{\text{mass of substance (g)}}{M_r}$$
4. Molar mass refers to the mass of one mole of a substance. It has the same value as A_r or M_r.

Test Yourself 9.2

Worked Example

A metal compound has the formula XCl_4 and a relative formula mass of 261. What is

a) the relative atomic mass of *X*?
b) metal *X*?

Thought Process

a) Let *M* be the A_r of metal *X*.
   M_r of $XCl_4 = M + 4 \times 35.5 = M + 142$
   Therefore, $M + 142 = 261$
   $M = 261 - 142 = 119$
b) Using the Periodic Table, the metal with an A_r of 119 is tin (Sn).

Answer

a) 119
b) Tin

Questions

1. A crucible contains 254 g of iodine I_2. How many moles of
   a) iodine molecules and
   b) iodine atoms are there in the crucible?
2. What is the mass of
   a) 0.25 mol of oxygen gas?
   b) 0.25 mol of nitrate ions (NO_3^-)?
3. Calculate
   a) the number of moles in 36 g of carbon.
   b) the mass of 0.4 mol of hydrogen sulphide (H_2S).
   c) the number of atoms in 6 g of magnesium.

9.5 | Percentage Composition of Compounds

Look at the pies in the picture. How can you tell which pie has more filling? You would need to examine each pie, for example, by breaking it into two. This is a form of analysis.

Chemists need to conduct analysis too, in order to find out how much of each element there is in a new compound. They do so by finding out the mass of each element in the compound. In this way, chemists know the percentage composition of a compound. Let us analyse the percentage composition of the compound hydrogen peroxide (H_2O_2).

How do we find the percentage composition of hydrogen peroxide?

In general, the percentage by mass of an element in a compound can be found using the formula:

Percentage by mass of an element in a compound

$$= \frac{A_r \text{ of element} \times \text{number of atoms in formula}}{\text{relative molecular mass } (M_r) \text{ of compound}} \times 100\%$$

The percentages (by mass) of hydrogen and oxygen present in hydrogen peroxide can be calculated from its chemical formula, H_2O_2.

Relative molecular mass of hydrogen peroxide (H_2O_2)

$= (2 \times 1) + (2 \times 16)$

$= 34$

Percentage of hydrogen in hydrogen peroxide

$$= \frac{A_r \text{ of hydrogen} \times \text{number of hydrogen atoms}}{\text{relative molecular mass of } H_2O_2} \times 100\%$$

$$= \frac{1 \times 2}{34} \times 100\%$$

$= \mathbf{5.9\%}$

Percentage of oxygen in hydrogen peroxide

$$= \frac{A_r \text{ of oxygen} \times \text{number of oxygen atoms}}{\text{relative molecular mass of } H_2O_2} \times 100\%$$

$$= \frac{16 \times 2}{34} \times 100\%$$

$= \mathbf{94.1\%}$

Example 1

Calculate the percentage of water in copper(II) sulphate crystals ($CuSO_4.5H_2O$).

Solution:

M_r of copper(II) sulphate crystals
= 64 + 32 + (4 × 16) + (5 × 18)
= 250

M_r of water
= (1 × 2) + 16
= 18

∴ Percentage of water

$$= \frac{M_r \text{ of water} \times \text{number of water molecules}}{M_r \text{ of copper(II) sulphate crystals}} \times 100\%$$

$$= \frac{18 \times 5}{250} \times 100\%$$

= **36%**

Key ideas

1. The percentage composition of a compound can be found given
   a) its formula,
   b) the relative atomic masses of elements in it.
2. Percentage composition by mass of an element in a compound

$$= \frac{A_r \text{ of element} \times \text{number of atoms of element in formula}}{M_r \text{ of compound}} \times 100\%$$

Test Yourself 9.3

Question

Calculate the following:

a) The percentage of nitrogen in potassium nitrate, KNO_3.
b) The percentage of chlorine in ammonium chloride, NH_4Cl.
c) The percentages of calcium and oxygen in calcium carbonate, $CaCO_3$.

9.6 | Finding the Formula of a Compound

Do you remember reading about the discovery of aniline in chapter 2? Chemists often discover useful new compounds while carrying out different tasks. In order to produce these useful products on a large scale, chemists need to know the chemical formulae of these compounds.

How can we work out the formula of a compound?

In the previous section, we saw that it was relatively easy to determine the percentage composition of a substance. This is because we already knew the formula of the compound. What if we did not know the formula of the compound?

We can conduct experiments to find out the formula of a compound.

1. First, we find out the mass of the reactants taking part in the reaction.
2. Next, we work out the relative numbers of moles of the reactants used.

We are then able to find the formula of the compound.

Experiment 1

To work out the formula of magnesium oxide produced by the combustion of magnesium

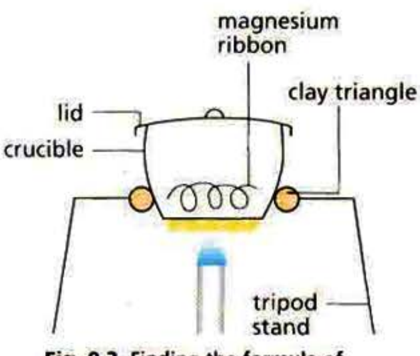


Fig. 9.2 Finding the formula of magnesium oxide

Procedure

1. Weigh a crucible together with the lid. Put a coil of magnesium ribbon in it and weigh again.
2. Put the lid on the crucible and heat the crucible gently (Fig. 9.2). When the magnesium catches fire (you will see a white glow through the crucible), heat it more strongly.
3. Use a pair of tongs to lift the lid slightly from time to time to allow air in. Quickly replace the lid to make sure that magnesium oxide formed does not escape.
4. When the burning is complete, allow the crucible to cool. Then weigh the crucible together with the lid and the magnesium oxide in it.

Sample results

Mass of crucible + lid = 26.52 g
Mass of crucible + lid + magnesium = 27.72 g
Mass of crucible + lid + magnesium oxide = 28.52 g

Calculations

Mass of magnesium = 27.72 − 26.52 = 1.20 g
Mass of magnesium oxide produced = 28.52 − 26.52 = 2.00 g
Mass of oxygen reacted = 2.00 − 1.20 = 0.80 g

Table 9.6 shows the masses of magnesium and oxygen that reacted in the experiment.

Element	Mg	O
Mass (obtained from experiment)	1.20 g	0.80 g
Relative atomic mass	24	16
Number of moles	$\frac{1.20}{24} = 0.05$	$\frac{0.80}{16} = 0.05$
Molar ratio (divide by smallest number from previous row)	$\frac{0.05}{0.05} = 1$	$\frac{0.05}{0.05} = 1$

smallest value in the 'number of moles' row

Table 9.6 Finding the number of moles of magnesium and oxygen that reacted

The simplest formula of magnesium oxide is thus **MgO**.

Empirical Formula

In experiment 1, we found that the formula MgO is the simplest formula that fits the experimental results. Other formulae such as Mg_2O_2 or Mg_3O_3 also fit the results since the ratios of magnesium to oxygen are the same. The formula MgO is called *the simplest formula or the empirical formula* of magnesium oxide. The **empirical formula** of a compound shows
- the types of elements present in it,
- the simplest ratio of the different types of atoms in it.

Chem-Aid

1. Remember: number of moles $= \frac{\text{mass}}{M_r}$
2. Notice that it is relative atomic mass, not relative molecular mass, that is used in deriving empirical formula.

Experiment 2

To determine the empirical formula of copper(II) oxide

Procedure

1. Weigh a porcelain boat. Put one spatula of copper(II) oxide (a black powder) in the boat and weigh again.
2. Put the porcelain boat containing copper(II) oxide into the middle of a Pyrex test tube with a small jet hole at its end.
3. Set up the apparatus that is shown in Fig. 9.3. Allow methane gas (from the gas tap) to pass through the apparatus for about 30 seconds to remove air.
4. Light the gas that escapes through the hole in the test tube. Then heat the copper(II) oxide until no further colour change is observed. The following reaction occurs during heating:

   copper(II) oxide + methane ⟶ copper + carbon dioxide + water

   Copper will remain as a brown residue in the porcelain boat.
5. Turn off the Bunsen burner but allow the methane gas to flow through the apparatus while the apparatus is cooling.
6. Turn off the gas supply when the apparatus is cool. Weigh the porcelain boat and its contents.

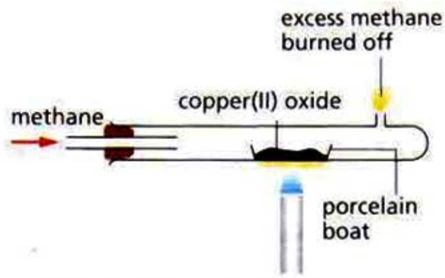


Fig. 9.3 Finding the formula of copper(II) oxide

Science Skills

a) Why is it necessary to remove all the air from the apparatus in Fig. 9.3?

b) Why is it necessary to allow the methane gas to flow while the apparatus is cooling?

Sample results

Mass of porcelain boat = 18.40 g
Mass of porcelain boat + copper(II) oxide = 35.89 g
Mass of porcelain boat + copper = 32.37 g

Calculations

Mass of copper = 32.37 − 18.40 = 13.97 g
Mass of oxygen = 35.89 − 32.37 = 3.52 g

Element	Cu	O
Collected mass (g)	13.97	3.52
Relative atomic mass	64	16
Number of moles	$\frac{13.97}{64} = 0.22$	$\frac{3.52}{16} = 0.22$
Molar ratio (divide by smallest number)	$\frac{0.22}{0.22} = 1$	$\frac{0.22}{0.22} = 1$

Table 9.7 Finding the number of moles of copper and oxygen that reacted

The empirical formula of copper(II) oxide is thus **CuO**.

Molecular Formula

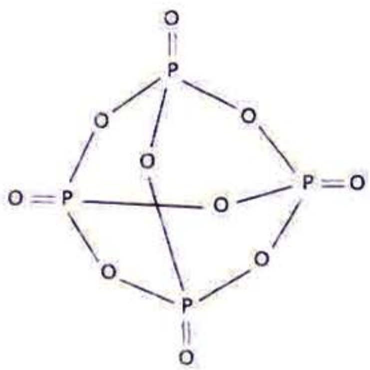


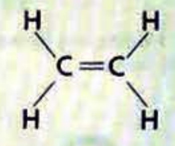
Fig. 9.4 Structure of phosphorus(V) oxide, P_4O_{10}

The empirical formula of phosphorus(V) oxide as determined by experiment is P_2O_5. Its actual formula is P_4O_{10} (Fig. 9.4). We call this the molecular formula. The **molecular formula** is *the formula that shows the exact number of atoms of each element in a molecule.*

What is the relationship between empirical formula and molecular formula?

For many compounds, such as water and ammonia, the empirical formula and molecular formula are the same. However, there are also many compounds (especially organic compounds) whose molecular formulae differ from their empirical formulae. The empirical formulae and molecular formulae of some common compounds are shown in Table 9.8.

Quick Check



What is the empirical formula of this compound?

Substance	Molecular formula	Empirical formula
water	H_2O	H_2O
ammonia	NH_3	NH_3
magnesium oxide	MgO	MgO
hydrogen peroxide	H_2O_2	HO
phosphorus(V) oxide	P_4O_{10}	P_2O_5
ethane	C_2H_6	CH_3

The molecular formula of a compound is a multiple of its empirical formula.

Table 9.8 Empirical formulae and molecular formulae of some common substances

Note that:

1. It is possible for different compounds to have the same empirical formula. For example, ethene (C_2H_4) and propene (C_3H_6) are two compounds with the same empirical formula, CH_2.

2. Where the empirical formula and molecular formula are different, the molecular formula is always a multiple of the empirical formula. For example, the molecular formula and empirical formula of phosphorus(V) oxide are P_4O_{10} and P_2O_5 respectively. The multiple is 2.

We can find the molecular formula of a substance if we know two things: the empirical formula and the relative molecular mass of the substance. They are related as follows:

If empirical formula = A_xB_y, molecular formula = $(A_xB_y)_n$ (where n = 1, 2, 3, etc.).
To find n,

$$n = \frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}}$$

Example 1

The empirical formula of ethane is CH_3. Given that the relative molecular mass of ethane is 30, what is its molecular formula?

Solution:

M_r of ethane from empirical formula = 12 + 3 = 15

$$\frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}} = \frac{30}{15} = 2$$

Hence, the molecular formula of ethane = $C_{1\times2}H_{3\times2}$ = $\mathbf{C_2H_6}$

Example 2

Compound *X* contains 40.0% carbon, 6.6% hydrogen and 53.3% oxygen. Its relative molecular mass is 180. What is the molecular formula of *X*?

Solution:

The percentage of each element is directly proportional to its mass in grams. Thus, the mass of each element in 100 g of compound is its percentage in the compound.

Element	C	H	O
Percentage in compound (%)	40.0	6.6	53.3
Relative atomic mass	12	1	16
Number of moles	$\frac{40.0}{12} = 3.3$	$\frac{6.6}{1} = 6.6$	$\frac{53.3}{16} = 3.3$
Molar ratio	$\frac{3.3}{3.3} = 1$	$\frac{6.6}{3.3} = 2$	$\frac{3.3}{3.3} = 1$

The empirical formula of X is thus CH_2O.

M_r of $CH_2O = 12 + (2 \times 1) + 16 = 30$

$$\frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}} = \frac{180}{30} = 6$$

Therefore, the molecular formula of $X = (CH_2O)_6 = \mathbf{C_6H_{12}O_6}$

Key ideas

1. The empirical formula shows the types of elements present in the simplest ratio in the compound.
2. The molecular formula shows the exact number of atoms of each element in a molecule.
3. $\dfrac{\text{relative molecular mass}}{M_r \text{ from empirical formula}} = n$ (where $n = 1, 2, 3$, etc.)

Test Yourself 9.4

Worked Example

Glycerol contains 39.1% carbon, 52.2% oxygen and the remainder is hydrogen. What is the molecular formula of glycerol? (1 mole of glycerol weighs 92.0 g.)

Answer

Element	C	H	O
Percentage in compound (%)	39.1	100 - 39.1 - 52.2 = 8.7	52.2
Relative atomic mass	12	1	16
Number of moles	$\frac{39.1}{12} = 3.3$	$\frac{8.7}{1} = 8.7$	$\frac{52.2}{16} = 3.3$
Molar ratio	$\frac{3.3}{3.3} = 1$	$\frac{8.7}{3.3} = 2.6$	$\frac{3.3}{3.3} = 1$
Simplest ratio	$1 \times 3 = \mathbf{3}$	$2.6 \times 3 = \mathbf{8}$	$1 \times 3 = \mathbf{3}$

Therefore, the empirical formula of glycerol is $C_3H_8O_3$.

The relative molecular mass from empirical formula
$= (3 \times 12) + (8 \times 1) + (3 \times 16)$
$= 92$

$$\frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}} = 1$$

Therefore, the molecular formula of glycerol $= (C_3H_8O_3)_1 = \mathbf{C_3H_8O_3}$

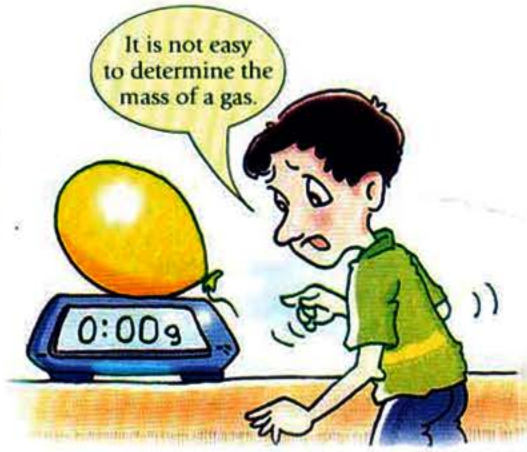
Questions

1. Caffeine is a compound found in coffee and tea. The percentage composition of caffeine is 49.5% carbon, 5.1% hydrogen, 16.5% oxygen and 28.9% nitrogen. The relative molecular mass of caffeine is 195. Determine
   a) the empirical formula and
   b) the molecular formula of caffeine.

2. The following results were obtained in an experiment to determine the formula of an oxide of silicon.
   Mass of crucible = 15.20 g
   Mass of crucible + silicon = 15.48 g
   Mass of crucible + oxide of silicon = 15.80 g
   a) Find the empirical formula of the oxide of silicon.
   b) If the M_r of the oxide of silicon is 60, what is its molecular formula?

3. A sample of hydrated copper(II) sulphate weighs 124.8 g. The sample has been determined to contain 31.8 g of copper(II) ions and 48.0 g of sulphate ions.
   a) How many molecules of water of crystallisation are present in the sample?
   b) Deduce the actual formula of hydrated copper(II) sulphate.

9.7 | Molar Gas Volume



You have learnt that the mass of one mole of a substance has a numerical value equal to its relative atomic mass or relative molecular mass. This property is applicable to solids, liquids and gases. However, it is easier to measure the volume of a gas than its mass, since gases weigh very little. Is there a way to relate moles to the volumes of gases?

> **Avogadro's Law** states that equal volumes of all gases, under the same conditions of temperature and pressure, contain the same number of molecules.

In fact, chemists have found by experiment that **one mole of any gas** occupies **24 dm^3** (24 000 cm^3) at room temperature and pressure (r.t.p.). This volume is called the **molar volume** of a gas.

Chem-Aid

r.t.p. or room temperature and pressure are often taken as the conditions of 25 °C and 1 atm.

This means that at r.t.p.
- 1 mol of oxygen occupies 24 dm^3,
- 1 mol of carbon dioxide occupies 24 dm^3,
- 2 mol of oxygen occupy 2 × 24 = 48 dm^3,
- 2 mol of carbon dioxide occupy 2 × 24 = 48 dm^3.

Does one mole of gas in a hot air balloon occupy 24 dm³?

How can we calculate the number of moles of a gas?

The number of moles of a gas can be measured in two ways.

1. Find the mass of the gas. Then use the following formula to calculate the number of moles of the gas.

   $$\text{Number of moles of gas} = \frac{\text{mass of gas in grams}}{M_r \text{ of gas}}$$

2. Find the volume of the gas. Then use this formula:

   $$\text{Number of moles of gas} = \frac{\text{volume of gas in cm}^3 \text{ at r.t.p.}}{24\,000 \text{ cm}^3}$$

   This formula can be rearranged to give

   Volume of gas (in cm^3) = number of moles × 24 000
   Volume of gas (in dm^3) = number of moles × 24

Example 1

What is the volume, in dm^3, of 8 g of oxygen gas (O_2) at r.t.p.?

Solution:

Relative molecular mass of oxygen = 2 × 16 = 32

$$\begin{aligned}\text{Volume of oxygen} &= \text{number of moles of oxygen} \times 24 \\ &= \frac{\text{mass of oxygen}}{M_r \text{ of oxygen}} \times 24 \\ &= \frac{8}{32} \times 24 \\ &= \mathbf{6\ dm^3}\end{aligned}$$

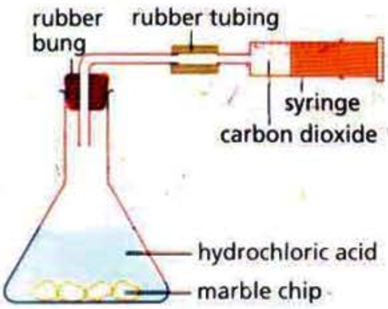


Fig. 9.5 Measuring the volume of carbon dioxide given off

Example 2

In an experiment, hydrochloric acid was reacted with calcium carbonate at room temperature and pressure (Fig. 9.5). 80 cm^3 of carbon dioxide was produced. Calculate the number of molecules of carbon dioxide given off.

Solution:

Number of moles of carbon dioxide given off

$$= \frac{80}{24\,000}$$

$= 3.33 \times 10^{-3}$ mol

Number of molecules of carbon dioxide given off

= number of moles × Avogadro's constant

$= 3.33 \times 10^{-3} \times 6 \times 10^{23}$

$= \mathbf{2.00 \times 10^{21}}$

Example 3

Calculate the mass of oxygen gas (O_2) in a room that measures 4 m high, 8 m wide and 10 m long. Assume that air contains 20% oxygen. ($1 \text{ m}^3 = 10^6 \text{ cm}^3$)

Solution:

Volume of air in the room $= 4 \times 8 \times 10$
$= 320 \text{ m}^3$
$= 320 \times 10^6 \text{ cm}^3$

Volume of oxygen in the room $= (320 \times 10^6) \times 20\%$
$= 64 \times 10^6 \text{ cm}^3$

Mass of oxygen $=$ number of moles of oxygen $\times M_r$

$$= \frac{\text{volume of oxygen}}{\text{molar volume}} \times 32$$

$$= \frac{64 \times 10^6 \text{ cm}^3}{24\,000 \text{ cm}^3} \times 32$$

$$= \mathbf{8.53 \times 10^4 \text{ g}}$$

Do the balloons of the same volume contain the same number of particles?

Look at the balloons on the right. They are of the same volume but do they contain the same number of particles?

Yes, they do! According to Avogadro's Law, each of these balloons contains the same number of gaseous particles since they have the same volume.

Balloons containing identical volumes of gas

Do the balloons of the same mass contain the same number of particles?

If the balloons each contained 0.18 g of a different gas (instead of having the same volume), they would *not* contain the same number of gaseous particles. The following calculation explains why.

Gas	Mass (g)	A_r or M_r	Number of moles $= \frac{\text{mass (g)}}{A_r \text{ or } M_r}$
helium (He)	0.18	4	$\frac{0.18}{4} = 0.045$
hydrogen (H_2)	0.18	2	$\frac{0.18}{2} = 0.090$
methane (CH_4)	0.18	16	$\frac{0.18}{16} = 0.011$

Table 9.9 Calculating the number of moles of different gases with equal mass

Number of helium atoms $= 0.045 \times 6 \times 10^{23} = 2.70 \times 10^{22}$
Number of hydrogen molecules $= 0.090 \times 6 \times 10^{23} = 5.40 \times 10^{22}$
Number of methane molecules $= 0.011 \times 6 \times 10^{23} = 6.60 \times 10^{21}$

Therefore, equal masses of different gases do *not* contain the same number of particles.

Key ideas

1. Avogadro's Law states that equal volumes of all gases under the same conditions of temperature and pressure contain the same number of particles.
2. The volume occupied by one mole of a gas is called its molar volume.
3. At room temperature and pressure, the molar volume of a gas is equal to 24 dm^3 or 24 000 cm^3.

Test Yourself 9.5

Worked Example

Which substance contains 6×10^{23} atoms at room temperature and pressure?

(A_r: N = 14; O = 16; Ne = 20; Mg = 24; Al = 27; Cl = 35.5)

A 12 000 cm^3 of neon
B 41 g of aluminium nitride (AlN)
C 12 dm^3 of oxygen
D 40 g of magnesium oxide (MgO)

Thought Process

Oxygen exists as diatomic molecules (O_2). 12 dm^3 of oxygen contain $\frac{12}{24}$ = 0.5 mol of oxygen molecules (O_2). Each molecule of oxygen contains two atoms of oxygen. Therefore, 0.5 mol of oxygen contains 1 mol of oxygen atoms. Hence, there are 6×10^{23} oxygen atoms.

Answer

C

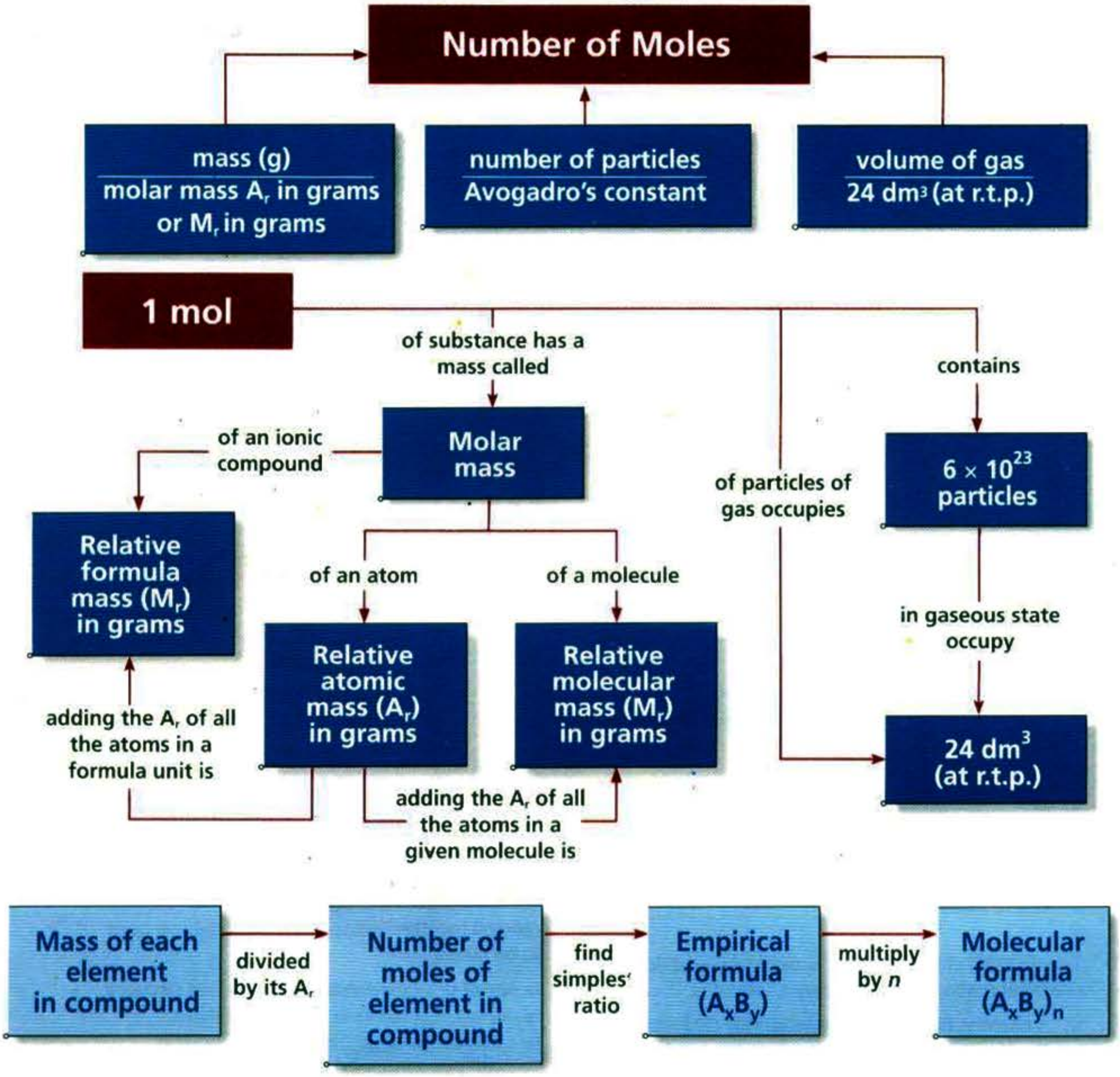
Questions

1. Calculate the number of
   a) sulphur atoms in 64 g of sulphur.
   b) magnesium atoms in 0.01 mol of magnesium.
   c) methane molecules in 112 cm^3 of methane at r.t.p.
   d) atoms in a drop of water weighing 0.5 g.

2. How many moles are present in the following volumes of gases (at r.t.p.)?
   a) 1.2 dm^3 of sulphur dioxide, SO_2
   b) 0.24 dm^3 of methane, CH_4
   c) 120 cm^3 of carbon dioxide, CO_2

3. 0.52 g of a metal (A_r = 65) reacts with 192 cm^3 of chlorine under room conditions to form a compound. What is the formula of the compound?

Concept Map



n = M_r divided by M_r of A_xB_y

Exercise 9

Foundation

1. Which step shows the correct way of calculating the relative formula mass of sodium carbonate crystals, $Na_2CO_3.10H_2O$?

   A $(2 \times 11) + 6 + (3 \times 8) + 10 \times 9$
   B $(2 \times 23) + (3 \times 16) + 10 \times 16$
   C $(2 \times 23) + 12 + (3 \times 16) + (10 \times 2) + (10 \times 16)$
   D $(2 \times 23) + 12 + (3 \times 16) + 10 + 160$

2. How many chlorine molecules are there in 71 g of gaseous chlorine?

   A $0.5 \times 6 \times 10^{23}$ B $2 \times 6 \times 10^{23}$
   C $35.5 \times 6 \times 10^{23}$ D 6×10^{23}

3. What amount, in moles, of calcium oxide is formed when 10.0 g of calcium is burnt in a plentiful supply of oxygen?

   A 0.125 B 0.25
   C 4 D 6

4. Find the empirical formula of a compound which contains 40% sulphur and 60% oxygen by mass.

   A S_2O B SO
   C SO_2 D SO_3

5. An ionic compound of the element *X* and chlorine has the formula *XCl*. The M_r of *XCl* is 74.5. What is the M_r of the oxide of *X*?

   A 55 B 71
   C 87 D 94

6. A sample of gas weighs 2.0 g and has a volume of 3 dm^3 at room conditions. What is the relative molecular mass of the gas?

7. Titanium(III) chloride, $TiCl_3$, is used to speed up chemical reactions in the plastics industry.

   a) A chemist has 38.6 g of titanium(III) chloride.
      i) How many moles of titanium(III) chloride does the chemist have?
      ii) What is the mass of titanium in this sample?

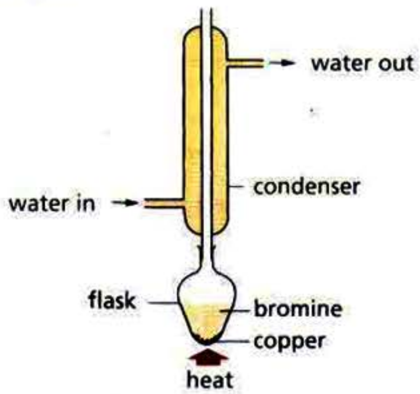
   b) How many chlorine atoms are there in 0.5 mol of titanium(III) chloride?

8. a) Calculate the percentage by mass of tin in tin(II) chloride ($SnCl_2$).

   b) Tin reacts with chlorine to form another chloride containing 45.6% of tin.
      i) What is the empirical formula of this chloride?
      ii) If the M_r of this chloride of tin is 261, what is the molecular formula of the chloride?

Challenge

A compound of copper and bromine was produced in a fume cupboard using the apparatus shown. Bromine has a melting point of −7 °C and a boiling point of 58 °C.



a) Suggest why
   i) the experiment was performed in a fume cupboard.
   ii) the above apparatus was used.

b) The flask was weighed empty. In order to work out the masses of copper and bromine that had reacted together, what **two** other readings need to be taken?

c) In one of the experiments, it was found that 64 g of copper reacted with 160 g of bromine.
   i) Calculate the empirical formula of the compound formed between copper and bromine.
   ii) What further information is required to find the molecular formula of the compound?

Chemistry Today

Did you know that cars are often fitted with air bags? The air bag is an extremely important device in protecting the life of a driver in case of a traffic accident.

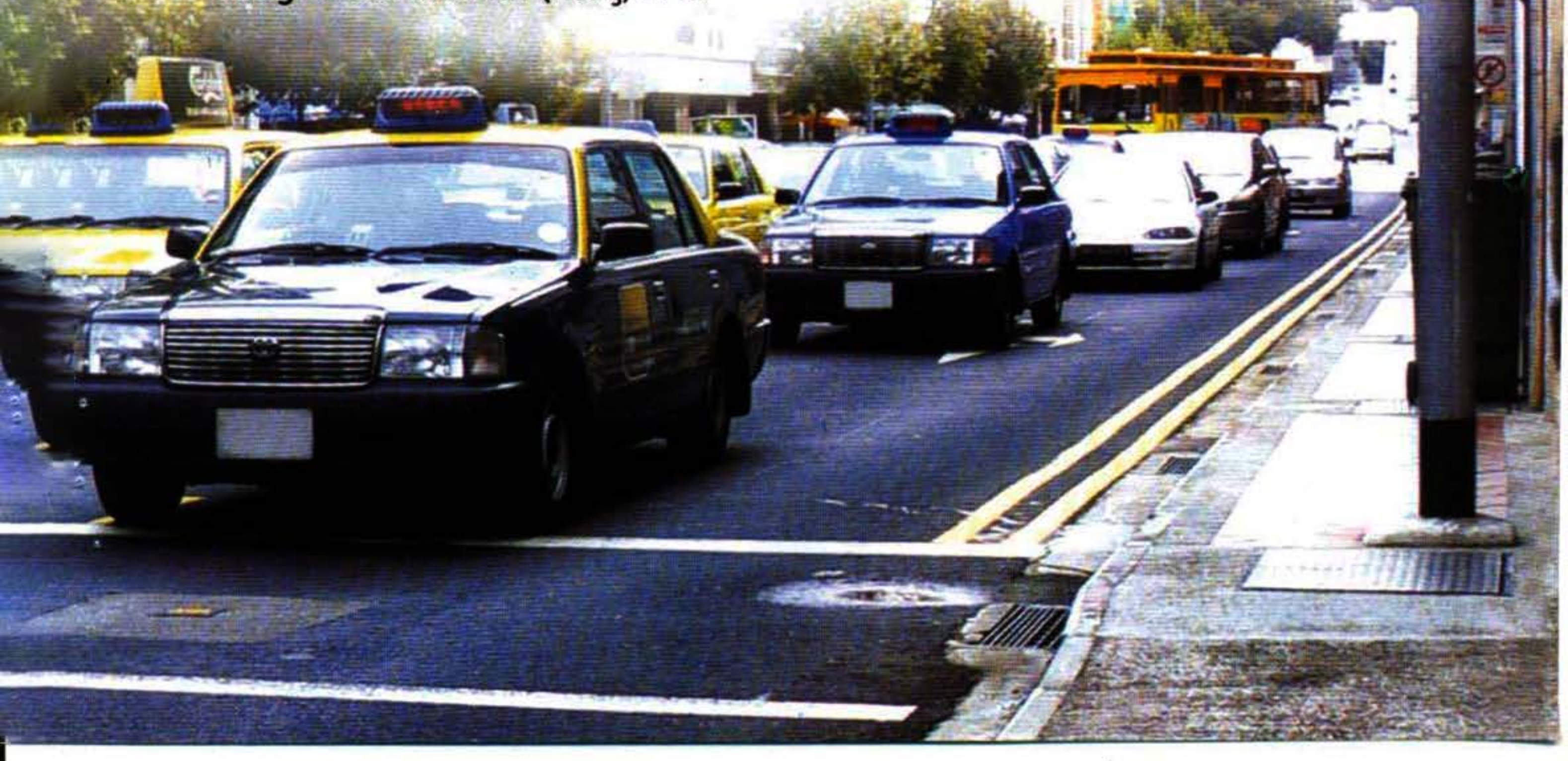
Under normal conditions, the air bag is deflated and stored in the steering wheel. In the event of a crash, it inflates to cushion the driver against the impact of the crash. This happens because of a chemical reaction that takes place in the air bag. The chemical equation is shown below.

$$2NaN_3 \longrightarrow 2Na + 3N_2$$

This reaction occurs very rapidly (which is crucial for saving lives) and only requires a very small amount of sodium azide (NaN_3).

CRITICAL THINKING

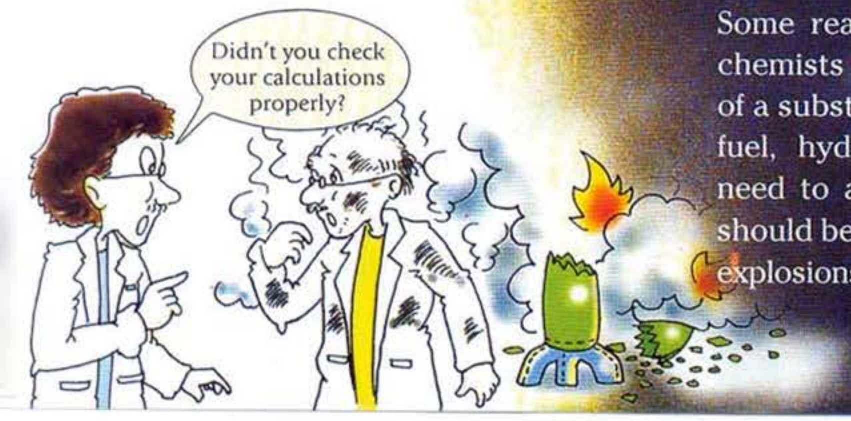
How big does an air bag become if there is **100 g of sodium azide (NaN_3) in it?**



Chapter **10**

Chemical Calculations

Chapter Outline



Chemical calculations are important when chemicals are produced on a large scale. Chemists use chemical calculations to predict the amounts of reactants needed and products expected from a chemical reaction.

Some reactions are also very dangerous and chemists need to calculate exactly how much of a substance is needed. For example, a rocket fuel, hydrazine, is very explosive. Chemists need to accurately calculate the amount that should be used in reactions as excess can lead to explosions.

10.1 | Calculations from Chemical Reactions

Writing a balanced chemical equation is often the first step in doing chemical calculations.

What does a chemical equation tell us?
A balanced chemical equation shows important facts about a reaction:
a) The reactants
b) The products
c) The ratio of the amounts (in moles) of the reactants and the products
d) The state of the reactants and products if indicated

Chem-Aid
An equation does not tell us
a) how fast or slow the reaction is,
b) whether the reaction gives out or takes in heat energy.

The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction is known as the **stoichiometry** of the reaction. Consider the following equation.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

We can read this equation as follows:

1 mol of magnesium	reacts with	2 mol of hydrochloric acid	to give	1 mol of magnesium chloride	and	1 mol of hydrogen

You learnt in the previous chapter that the mass of one mole of a substance is equal to its M_r in grams. Hence, we can deduce that

24 g of magnesium	reacts with	73 g of hydrochloric acid	to give	95 g of magnesium chloride	and	2 g of hydrogen

In other words, using a balanced equation, we can calculate the mass of any reactant consumed (used up) or product formed in a reaction.

Sodium hydrogencarbonate is used to make baking powder.

Example 1
Calculate the mass of the solid obtained when 16.8 g of sodium hydrogencarbonate is heated strongly until there is no further change. The equation for the reaction is

sodium hydrogencarbonate
$\longrightarrow$ sodium carbonate + carbon dioxide + water — **Word equation**

$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$ — **Chemical equation**

Solution:
M_r of $NaHCO_3$ = 23 + 1 + 12 + (3 × 16) = 84
M_r of Na_2CO_3 = (2 × 23) + 12 + (3 × 16) = 106

Number of moles of $NaHCO_3$ used $= \frac{16.8}{84} = 0.2$ mol

According to the equation,
2 mol of $NaHCO_3$ produce 1 mol of Na_2CO_3.

Number of moles of Na_2CO_3 obtained $= \frac{0.2 \times 1}{2} = 0.1$ mol

∴ Mass of solid obtained
= number of moles of Na_2CO_3 × M_r of Na_2CO_3
= 0.1 × 106
= **10.6 g**

Example 2
63.0 dm^3 of carbon monoxide, measured at r.t.p., was used to react with iron(III) oxide. What mass of iron was produced at the end of the reaction? The equation for the reaction is

Word equation iron(III) oxide + carbon monoxide ⟶ iron + carbon dioxide

Chemical equation $Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$

Solution:
Number of moles of carbon monoxide used $= \frac{63.0}{24} = 2.625$ mol

From the equation,
3 mol of carbon monoxide produce 2 mol of iron.

Number of moles of iron produced in the reaction $= \frac{2.625 \times 2}{3}$
= 1.75 mol

A_r of Fe = 56
∴ Mass of iron produced = number of moles of Fe × A_r of Fe
= 1.75 × 56
= **98.0 g**

Chem-Aid
Number of moles of gas at r.t.p.
$= \frac{\text{actual volume (dm}^3)}{24 \text{ dm}^3}$

Link
Example 2 is one of the reactions that occur in a blast furnace in the extraction of iron. Find out more about extraction of iron in chapter 14.

10.2 | The Volumes of Reacting Gases

Ammonia gas (NH_3) is made on a large scale by reacting hydrogen gas with nitrogen gas. Many other industrial processes also involve reacting gases.

How do we do calculations involving gases?
In chapter 9, we learnt that one mole of any gas occupies 24 dm^3 at room temperature and pressure. This means that *the volume of gas is proportional to the number of moles of the gas, and vice versa.* Thus, for chemical calculations that involve gases, we can change mole of gas to volume of gas.

Consider the following equation.

$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

We can interpret it as

1 mol of nitrogen	reacts with	3 mol of hydrogen	to form	2 mol of ammonia

OR

1 volume of nitrogen	reacts with	3 volumes of hydrogen	to form	2 volumes of ammonia

Example 1
Refer to the equation below.

nitrogen + hydrogen $\longrightarrow$ ammonia **Word equation**

$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ **Chemical equation**

a) Calculate the volume of ammonia gas produced when 20 cm^3 of nitrogen gas was reacted with an excess of hydrogen gas.
b) What is the volume of hydrogen required for this reaction?

Solution:
a) According to the equation,
1 mol of nitrogen produces 2 mol of ammonia,
i.e. 1 volume of nitrogen produces 2 volumes of ammonia.
$\therefore$ Volume of ammonia produced = 2×20 = **40 cm^3**

b) Also, 1 volume of nitrogen reacts with 3 volumes of hydrogen.
$\therefore$ Volume of hydrogen required = 3×20 = **60 cm^3**

Example 2
At room temperature and pressure, 25 cm^3 of butane (C_4H_{10}) exploded after it was mixed with an excess of oxygen. Calculate
a) the volume of carbon dioxide produced, and
b) the volume of oxygen required for the reaction.

butane + oxygen $\longrightarrow$ carbon dioxide + water **Word equation**

$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(l)$ **Chemical equation**

Solution:

a) According to the equation,
2 volumes of butane produce 8 volumes of carbon dioxide.

$\therefore$ Volume of carbon dioxide produced = $\frac{8}{2} \times 25$ = **100 cm³**

b) 2 volumes of butane react with 13 volumes of oxygen.

$\therefore$ Volume of oxygen required = $\frac{13}{2} \times 25$ = **162.5 cm³**

10.3 | Limiting Reactants

Is it possible to clean your entire home with this amount of soap?

It is quite impossible to clean your entire home with a small amount of soap. The reason is simply because there is not enough soap to remove all the dirt. As there is insufficient soap, chemists would say that soap is the **limiting reactant**.

For any chemical reaction, it is possible to calculate the exact quantities of reactants that are required and products that are formed from a balanced chemical equation. Ideally, reactions should be carried out using exact quantities of reactants to reduce wastage. However, many reactions are carried out using an excess amount of one reactant. This ensures that the more expensive reactant is completely used up. To do so, we make use of the idea of limiting reactants.

What limits the amount of products formed?
Consider the reaction between hydrogen and chlorine to form hydrogen chloride. The chemical equation for the reaction is shown below.

Word equation hydrogen + chlorine ⟶ hydrogen chloride

Chemical equation $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

From the equation, we can see that one mole of hydrogen reacts with one mole of chlorine to produce two moles of hydrogen chloride. Table 10.1 shows the results of three experiments obtained by using different molar ratios of the reactants, hydrogen and chlorine.

In experiment A, the ratio of the number of moles of reactant molecules used is the same as those given in the balanced chemical equation — one mole of hydrogen reacts with one mole of chlorine to give two moles of hydrogen chloride.

In experiments B and C, the ratio of the number of moles of reactant molecules used is different from those given in the balanced chemical equation.

Experiment	Amount of reactants used		Amount of products formed		
	Hydrogen (H_2)	Chlorine (Cl_2)	Hydrogen chloride (HCl)	Unreacted hydrogen	Unreacted chlorine
A	1	1	2	0	0
	H H	Cl Cl	H Cl H Cl		
B	2	1	2	1	0
	H H H H	Cl Cl	H Cl H Cl	H H	
C	1	2	2	0	1
	H H	Cl Cl Cl Cl	H Cl H Cl		Cl Cl

All the reactants are used up.

Excess hydrogen is used. Hence, some hydrogen is unreacted. **In this case, chlorine is the limiting reactant.**

Excess chlorine is used. Hence, some chlorine is unreacted. **In this case, hydrogen is the limiting reactant.**

Table 10.1 Results of experiments using different molar ratios of reactants (hydrogen and chlorine)

The reactant that is completely used up in a reaction is known as the **limiting reactant**. It is called the limiting reactant because it *determines or limits the amount of products formed.* The reactants that are not used up are called the **excess reactants**.

What is the effect of limiting reactants?

The quantity of products formed in a reaction is always determined by the quantity of the limiting reactant. For example, in experiment B, chlorine acts as the limiting reactant because it limits the amount of hydrogen chloride produced to two moles, regardless of the amount of hydrogen used. Once all the chlorine has been used up, the reaction stops. Hydrogen is the excess reactant and some is left over at the end of the reaction.

In experiment C, hydrogen acts as the limiting reactant because it limits the amount of hydrogen chloride produced to two moles, regardless of the amount of chlorine used. Once all the hydrogen has been used up, the reaction stops. Chlorine is the excess reactant and some is left over at the end of the reaction.

Example 1

Zinc reacts with hydrochloric acid according to the equation below.

Word equation zinc + hydrochloric acid ⟶ zinc chloride + hydrogen

Chemical equation $Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

If 0.05 mol of zinc was added to 0.075 mol of hydrochloric acid,

a) identify the limiting reactant.
b) calculate the amount (in moles) of the excess reactant that remained unreacted.

Solution:

a) According to the equation,
1 mol of Zn will react with 2 mol of HC*l*,
therefore 0.05 mol of Zn will react with 0.10 mol of HC*l*.

0.075 mol of HCl was used and this will react with 0.0375 mol of Zn. Since 0.05 mol of Zn were used, the zinc must be in excess and **HC*l* is the limiting reactant.**

b) Amount of zinc which remained unreacted = 0.05 − 0.0375
= **0.0125 mol**

Example 2

Word equation ethene + oxygen ⟶ carbon dioxide + water

Chemical equation $C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l)$

Ethene, C_2H_4, burns in oxygen according to the above equation. In an experiment, 10 cm^3 of ethene was burnt in 50 cm^3 of oxygen.

a) Which gas was supplied in excess? Calculate the volume of the excess gas remaining at the end of the reaction.
b) Calculate the volume of carbon dioxide produced.

Solution:

a) According to the equation,
1 volume of ethene reacts with 3 volumes of oxygen.

$\therefore$ Volume of oxygen used = $\frac{3}{1} \times 10 = 30\ cm^3$

However, 50 cm^3 of oxygen was used.

$\therefore$ **Oxygen gas was in excess.**

Hence, volume of O_2 remaining = initial volume − volume used
= 50 − 30
= **20 cm^3**

b) 1 volume of ethene produces 2 volumes of carbon dioxide.

$\therefore$ Volume of carbon dioxide produced = $\frac{2}{1} \times 10$
= **20 cm^3**

Ethene is an important material for making plastic toys.

Why is it important to identify the limiting reactant?

In the chemical industry, large amounts of chemicals are required to manufacture a particular product. To get the maximum yield of a product at the minimum cost, we need to know the limiting reactant. We will generally choose the most expensive reactant to be the limiting reactant, and use excess amounts of the other reactants in a reaction. This ensures all the expensive reactant is used up.

In many industrial reactions, excess reactants are recycled as far as possible in order to reduce production costs.

Key ideas

1. **A balanced chemical equation is used to calculate the amounts of reactants that are consumed and the products that are formed.**
2. **The limiting reactant determines the amount of products formed.**

Test Yourself 10.1

Worked Example

0.024 g of magnesium was reacted with excess hydrochloric acid at room temperature and pressure.

a) Write the equation for this reaction.
b) How many moles of magnesium were used?
c) What volume of hydrogen is produced in this reaction? (1 mol of gas at r.t.p. occupies 24 dm^3.)

Answer

a) $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$

b) Number of moles of magnesium used $= \frac{0.024}{24} = 0.001$ mol

c) From the equation, 1 mol of magnesium produces 1 mol of hydrogen. Therefore, 0.001 mol of magnesium produces 0.001 mol of hydrogen.

Volume of hydrogen produced $= 0.001 \times 24 = 0.024$ dm^3

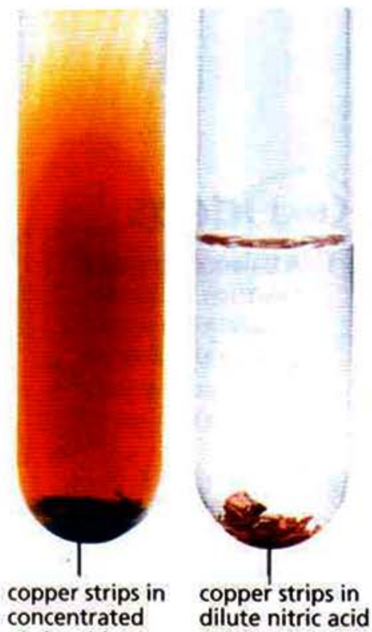
Questions

1. a) Write the balanced equation for the reaction between sulphur and oxygen to form sulphur dioxide.
   b) In one process, 4.0 g of sulphur was burnt in 48.0 dm^3 of oxygen measured at r.t.p.
      i) What is the limiting reactant in this reaction?
      ii) Calculate the volume of sulphur dioxide formed at r.t.p.

2. Phosphorus(III) chloride, PCl_3, is used in the manufacture of pesticides. It is made by the direct combination of phosphorus and chlorine.

   $P_4(s) + 6Cl_2(g) \longrightarrow 4PCl_3(l)$

   Calculate the mass of PCl_3 produced when 125 g of P_4 was allowed to react with 323 g of chlorine gas.



The concentration of an acid can affect a reaction. Copper reacts with concentrated nitric acid, but not with dilute nitric acid.

10.4 | The Concentration of a Solution

Chemical reactions often involve solutions of substances in water. It is therefore important to know the concentration of a solution, especially in cases where the concentrations of the reactants can greatly affect the rate of the chemical reaction or the type of product obtained.

How do we determine the concentration of a solution?

The **concentration** of a solution is given by *the amount of a solute dissolved in a unit volume of the solution*. We can express concentration in terms of grams of solute per dm^3 (g/dm^3) of solution. 1 dm^3 is equivalent to 1000 cm^3.

$$\text{Concentration } (g/dm^3) = \frac{\text{mass of solute in grams}}{\text{volume of solution in } dm^3}$$

By rearranging the equation,

$$\text{Mass of solute (g)} = \text{concentration } (g/dm^3) \times \text{volume } (dm^3)$$

Example 1

A solution of glucose contains 0.45 g of glucose in 75 cm^3 of solution. What is the concentration of the glucose solution in g/dm^3?

Solution:

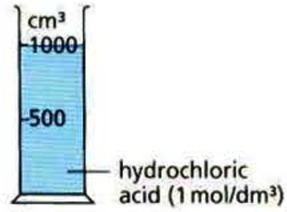
$$\text{Volume of glucose solution} = \frac{75}{1000}\ dm^3 = \mathbf{0.075\ dm^3}$$

$$\text{Concentration of glucose solution} = \frac{\text{mass (g)}}{\text{volume } (dm^3)} = \frac{0.45}{0.075} = \mathbf{6.00\ g/dm^3}$$

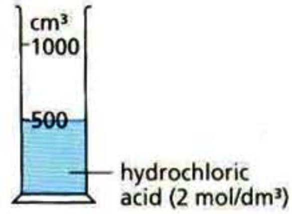
There are times when glucose solution must be introduced into the bodies of patients. Why is it dangerous for this solution to be highly concentrated?

Molar Concentration

The concentration of a solution can also be expressed in **mol/dm^3**. When *concentration is expressed in mol/dm^3*, it is called **molar concentration**. A solution of concentration 1 mol/dm^3 has one mole of solute dissolved in 1 dm^3 of solution.



This measuring cylinder contains one mole of HCl.



This measuring cylinder also contains one mole of HCl.

Fig. 10.1 Two measuring cylinders containing equal amounts of HCl

We can calculate the concentration in mol/dm^3 of a solution by using the formula

$$\text{Concentration (mol/dm}^3) = \frac{\text{number of moles of solute}}{\text{volume of solution in dm}^3}$$

OR

$$\text{Concentration (mol/dm}^3) = \frac{\text{concentration (g/dm}^3)}{M_r}$$

A solution of sodium hydroxide was prepared by dissolving 3.5 g of sodium hydroxide in distilled water and making the volume up to 100 cm^3. What is the concentration of the sodium hydroxide solution in mol/dm^3?

M_r of sodium hydroxide (NaOH) = 23 + 16 + 1 = 40

Number of moles of NaOH used $= \frac{\text{mass}}{M_r} = \frac{3.5}{40} = 0.0875$ mol

Volume of NaOH = 100 cm^3

$= \frac{100}{1000}$

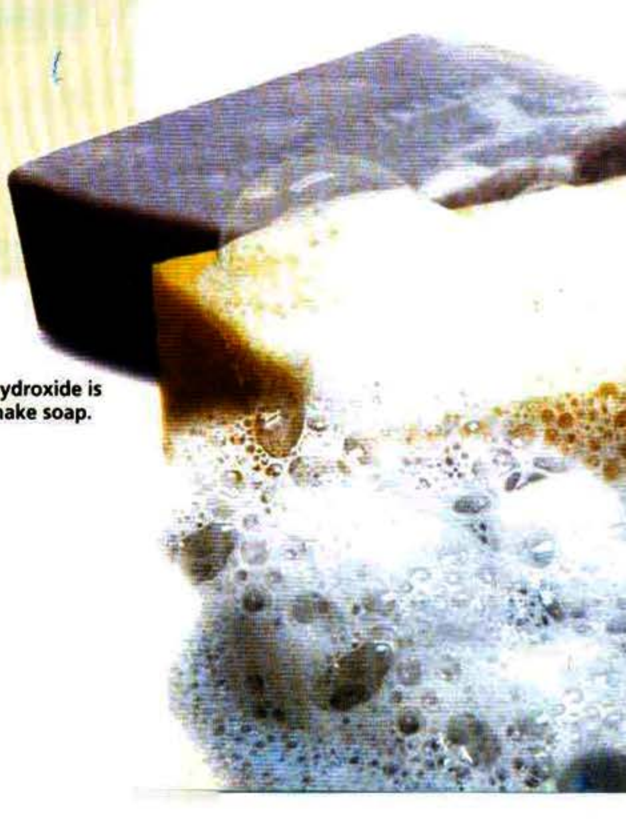
$= 0.10$ dm^3

Concentration of NaOH solution in mol/dm^3

$= \frac{\text{number of moles of NaOH}}{\text{volume of solution (dm}^3)}$

$= \frac{0.0875}{0.10}$

= **0.875 mol/dm^3**



Sodium hydroxide is used to make soap.

Example 2
The concentration of a bottle of hydrochloric acid is 73 g/dm^3. How many moles of HCl are there in 250 cm^3 of this acid?

Solution:
M_r of HCl = 1 + 35.5 = 36.5

Moles of HCl in 73 g $= \frac{73}{36.5} = 2$ mol

$\therefore$ 1000 cm^3 of the HCl solution contains 2 mol of HCl.

250 cm^3 will contain $2 \times \frac{250}{1000}$ = **0.5 mol of HCl**

Example 3
The concentration of a solution of HCl is 30 g/dm^3. What is its concentration in mol/dm^3?

Solution:
Concentration of HCl in mol/dm^3

$$= \frac{\text{concentration in g/dm}^3}{M_r \text{ of HCl}}$$

$$= \frac{30}{1 + 35.5}$$

= **0.822 mol/dm^3**

Key ideas

1. Concentration (g/dm^3) $= \frac{\text{mass of solute in grams}}{\text{volume of solution in dm}^3}$
2. Concentration (mol/dm^3) $= \frac{\text{number of moles of solute}}{\text{volume of solution in dm}^3}$

   $= \frac{\text{concentration (g/dm}^3)}{M_r \text{ of solute}}$
3. Number of moles of solute = concentration (mol/dm^3) × volume (dm^3)

Sulphuric acid is one of the most important chemicals. It can be used to make dyes, paints and even fertilisers.

Test Yourself 10.2

Worked Example

The concentration of dilute sulphuric acid found in school laboratories is often 2 mol/dm^3. What is the mass of sulphuric acid in 250 cm^3 of this acid?

Answer

M_r of H_2SO_4 = (2 × 1) + (1 × 32) + (4 × 16) = 98

1000 cm^3 of the acid contains 2 mol (given) of H_2SO_4.

250 cm^3 will contain $2 \times \frac{250}{1000} = 0.5$ mol of H_2SO_4

Mass of sulphuric acid = 0.5 × 98 = **49 g**

Questions

1. 400 cm^3 of 0.5 mol/dm^3 hydrochloric acid was reacted with 5 g of magnesium oxide. Calculate the mass of reacted oxide.

2. 14.0 g of potassium hydroxide was dissolved to form 100 cm^3 of an aqueous solution. What was the concentration of this solution in mol/dm^3 ?

3. A patient took some tablets containing magnesium oxide to relieve his gastric pain. Each tablet contains 3.0 g of magnesium oxide. The pain was caused by an excess amount of hydrochloric acid in his stomach. If his stomach contains the equivalent of 100 cm^3 of excess hydrochloric acid of concentration 3.0 mol/dm^3, how many tablets should he take to relieve the pain?

10.5 | Volumetric Anslysis

Many products that we use contain chemicals dissolved in water. These chemicals can be harmful when they are present in large quantities. For example, floor-cleaning products containing ammonia are harmful if the concentration of ammonia exceeds a certain level. Hence, a chemist needs to perform experiments to check the concentration of these substances. To do so, a chemist performs **volumetric analysis**.

How can we determine the concentration of a solution by volumetric analysis?

To do volumetric analysis, we use a method called **titration**. *In titration experiments, we determine the volume of a solution (chemical reagent) required to completely react with a known volume of another solution.* By measuring volumes and doing some calculations, we can determine the concentration of a solution.

Titration experiments that involve the use of an acid, e.g. HC*l*, and a base, e.g. NaOH, are called an acid-base titration.

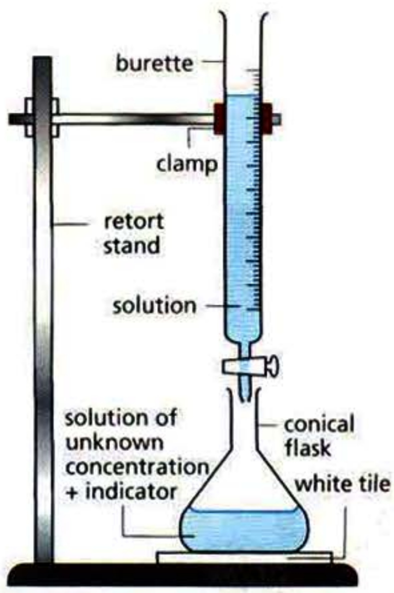


Fig. 10.2 Apparatus for titration experiments

Chem-Aid

A titrant is a solution placed in the burette in a titration experiment.

Indicator in acidic condition | Indicator at end-point | Indicator in alkaline condition

An indicator, such as methyl orange shown above, changes colour during a titration experiment. Find out more about indicators in chapter 11.

How do we perform titration?

1. The titrant is placed in the burette. The solution of unknown concentration is introduced into a conical flask using a pipette. One or two drops of indicator are then added to the conical flask.
2. The titrant is allowed to react completely with a known volume (usually 20 cm^3 or 25 cm^3) of the solution of unknown concentration.
3. Titration is stopped at the **end-point**. The end-point is reached when the indicator permanently changes colour. The volume of titrant used is noted.
4. The concentration of the unknown solution can then be calculated.

Example 1

A household ammonia solution was analysed to determine its ammonia content. 25 cm^3 of the ammonia solution required 21.9 cm^3 of 0.11 mol/dm^3 sulphuric acid to achieve the end-point of titration. Calculate the concentration of ammonia, in g/dm^3, in the household ammonia solution.

Solution:

Word equation sulphuric acid + ammonia $\longrightarrow$ ammonium sulphate

Chemical equation $H_2SO_4(aq) + 2NH_3(aq) \longrightarrow (NH_4)_2SO_4(aq)$

Number of moles of sulphuric acid used $= 0.11 \times \frac{21.9}{1000}$

$= 2.409 \times 10^{-3}$ mol

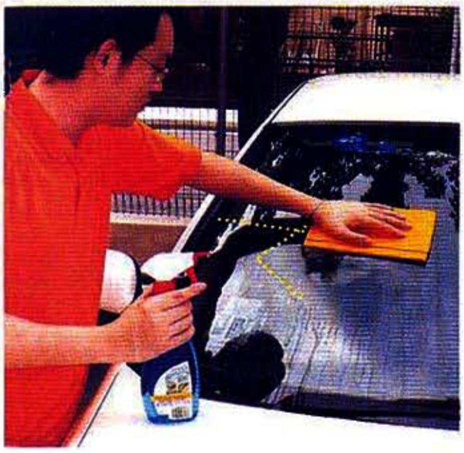
From the equation,
1 mol of sulphuric acid reacts with 2 mol of ammonia solution.

$\therefore$ Number of moles of ammonia in 25 cm^3 $= 2 \times 2.409 \times 10^{-3}$
$= 4.818 \times 10^{-3}$ mol

$\therefore$ Number of moles of ammonia in 1000 cm^3 $= 4.818 \times 10^{-3} \times \frac{1000}{25}$
$= 0.1927$ mol

M_r of $NH_3 = 14 + (3 \times 1) = 17$

$\therefore$ Concentration of ammonia in g/dm^3 $= 0.1927 \times 17$
$=$ **3.28 g/dm^3**



Ammonia is used to make cleansers.

Example 2

16.6 g of a metal carbonate, M_2CO_3, was made up to 1000 cm^3 of aqueous solution. 25 cm^3 of this solution required 30 cm^3 of 0.20 mol/dm^3 HC*l* for complete reaction.

a) Calculate the number of moles of HC*l* used in this reaction.
b) Write the equation for the reaction between M_2CO_3 and HC*l*.
c) Calculate the number of moles of M_2CO_3 present in
  i) 25 cm^3 of solution.
  ii) 1 dm^3 of solution.
d) Calculate
  i) the relative atomic mass of *M*.
  ii) the relative formula mass (M_r) of M_2CO_3.
e) Identify the metal *M*.

Solution:

a) Number of moles of HC*l* used = $0.2 \times \frac{30}{1000}$ = **0.006 mol**

b) The equation for the reaction is

$$M_2CO_3(aq) + 2HCl(aq) \longrightarrow 2MCl(aq) + H_2O(l) + CO_2(g)$$

c) i) From the equation,
1 mol of M_2CO_3 reacts with 2 mol of HC*l*.

Since the number of moles of HC*l* used was 0.006,
the number of moles of M_2CO_3 in 25 cm^3 = **0.003 mol**

ii) ∴ There is $0.003 \times \frac{1000}{25}$ in 1000 cm^3 (1 dm^3)
= **0.12 mol of M_2CO_3** in 1 dm^3 of the solution.

d) i) Let the relative atomic mass of *M* be Q.
The relative formula mass of M_2CO_3
$= (2 \times Q) + (1 \times 12) + (3 \times 16)$
$= 2Q + 60$

Number of moles of M_2CO_3 in 16.6 g = $\frac{16.6}{2Q + 60}$

∴ From (c)(ii),

$0.12 = \frac{16.6}{2Q + 60}$

$0.24Q + 7.2 = 16.6$

$Q =$ **39.2**

ii) Relative formula mass = $2 \times 39.2 + 60$ = **138.4**

e) From (d)(i), the relative atomic mass of *M* is 39.2. By checking the Periodic Table, we identify *M* as **potassium**.

Link

The results obtained from titration can be used to identify an unknown substance, as shown here in example 2. Titration can also be used to make salts in the laboratory. Find out more in Chapter 12.

Key ideas

Volumetric analysis is a technique used to determine the volumes of solutions that react together. In volumetric analysis, we perform titration.

Test Yourself 10.3

Worked Example

What volume (in cm^3) of 0.100 mol/dm^3 copper(II) sulphate solution is required to react completely with a solution containing 0.0250 mol of sodium hydroxide?

Answer

The equation for the reaction is

$$CuSO_4(aq) + 2NaOH(aq) \longrightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

From the equation,
1 mol of copper(II) sulphate reacts with 2 mol of sodium hydroxide.
0.0125 mol of copper(II) sulphate reacts with 0.0250 mol of sodium hydroxide.

There is 0.100 mol of copper(II) sulphate in 1000 cm^3.
$\therefore$ Volume for 0.0125 mol of copper(II) sulphate will be

$$1000 \times \frac{0.0125}{0.100} = \mathbf{125\ cm^3}$$

Question

Potassium hydroxide reacts with sulphuric acid according to the equation below.

$$2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$$

Calculate the minimum volume of 0.1 mol/dm^3 sulphuric acid that is required to completely react with 25.0 cm^3 of 0.4 mol/dm^3 potassium hydroxide.

10.6 | Percentage Yield and Percentage Purity

In the chemical calculations that we have done so far, we assumed that all the reactants are converted to products. The amount of products formed in a reaction is known as the **yield**. The **theoretical yield** of a reaction is the *calculated amount of products that would be obtained if the reaction is completed.*

In practice, few reactions occur this way. Most reactions do not go to completion. *The amount of pure products that is actually produced in the experiment* is called the **actual yield**. The actual yield is almost always *less than* the theoretical yield. (It can never be greater than the theoretical yield!)

Chem-Aid

In practice, a *side reaction* or a *reversible reaction* may occur in some reactions. A **side reaction** is a reaction that occurs in a way different from that desired. You will learn about reversible reactions in chapter 19.

The **percentage yield** shows the relationship between actual yield and theoretical yield.

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Often, the percentage yield is less than 100% because the reactants are not pure. The more impure the reactants, the lower the actual yield of the product.

We can calculate the percentage purity of a substance using this formula:

$$\text{Percentage purity} = \frac{\text{amount of pure substance}}{\text{amount of substance used in reaction}} \times 100\%$$

Example 1

When 1.92 g of magnesium was heated in excess oxygen, 3.0 g of magnesium oxide was obtained. Calculate the percentage yield of magnesium oxide.

Solution:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Molar ratio of Mg : MgO = 2 : 2 = 1 : 1

This means that 24 g of Mg will produce 40 g of MgO.

Therefore, 1.92 g of Mg should produce $1.92 \times \frac{40}{24} = 3.2$ g of MgO.

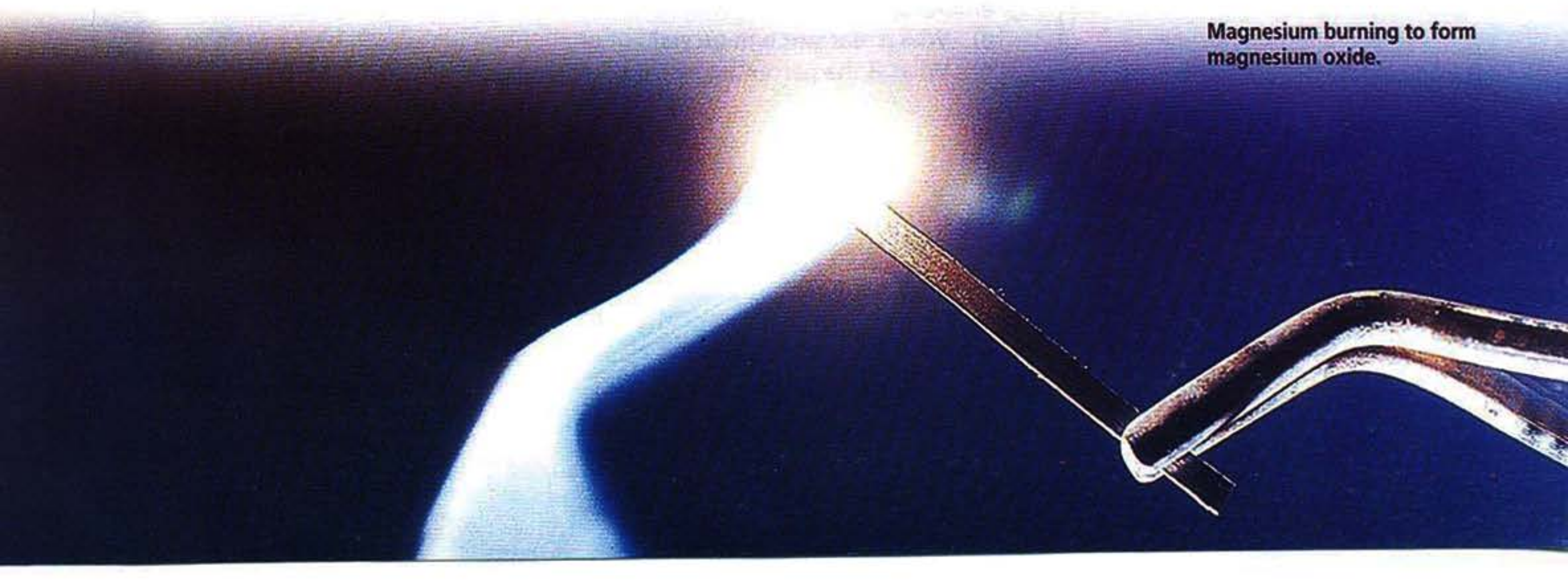
Actual amount of magnesium oxide obtained was 3.0 g.

Percentage yield of magnesium oxide

$= \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$

$= \frac{3.0}{3.2} \times 100\ \%$

$=$ **93.8%**



Magnesium burning to form magnesium oxide.

Key ideas

1. The yield of a reaction is usually not 100%. This is because a reaction seldom go to completion.
2. Percentage yield $= \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$
3. Percentage purity $= \dfrac{\text{amount of pure substance}}{\text{amount of substance used in reaction}} \times 100\%$

Test Yourself 10.4

Worked Example

Nitrogen and hydrogen react to form ammonia.

$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

12 dm^3 of hydrogen reacted with excess nitrogen to form 2 dm^3 of ammonia. What is the percentage yield of ammonia at room temperature and pressure? (1 mol of gas occupies 24 dm^3 at room temperature and pressure.)

Answer

From the equation, 3 mol of hydrogen give 2 mol of ammonia.

Moles of hydrogen used $= \dfrac{12}{24} = 0.5$ mol

Theoretical moles of ammonia formed $= 0.5 \times \dfrac{2}{3} = 0.33$ mol

Actual volume of ammonia $= 2\ dm^3$

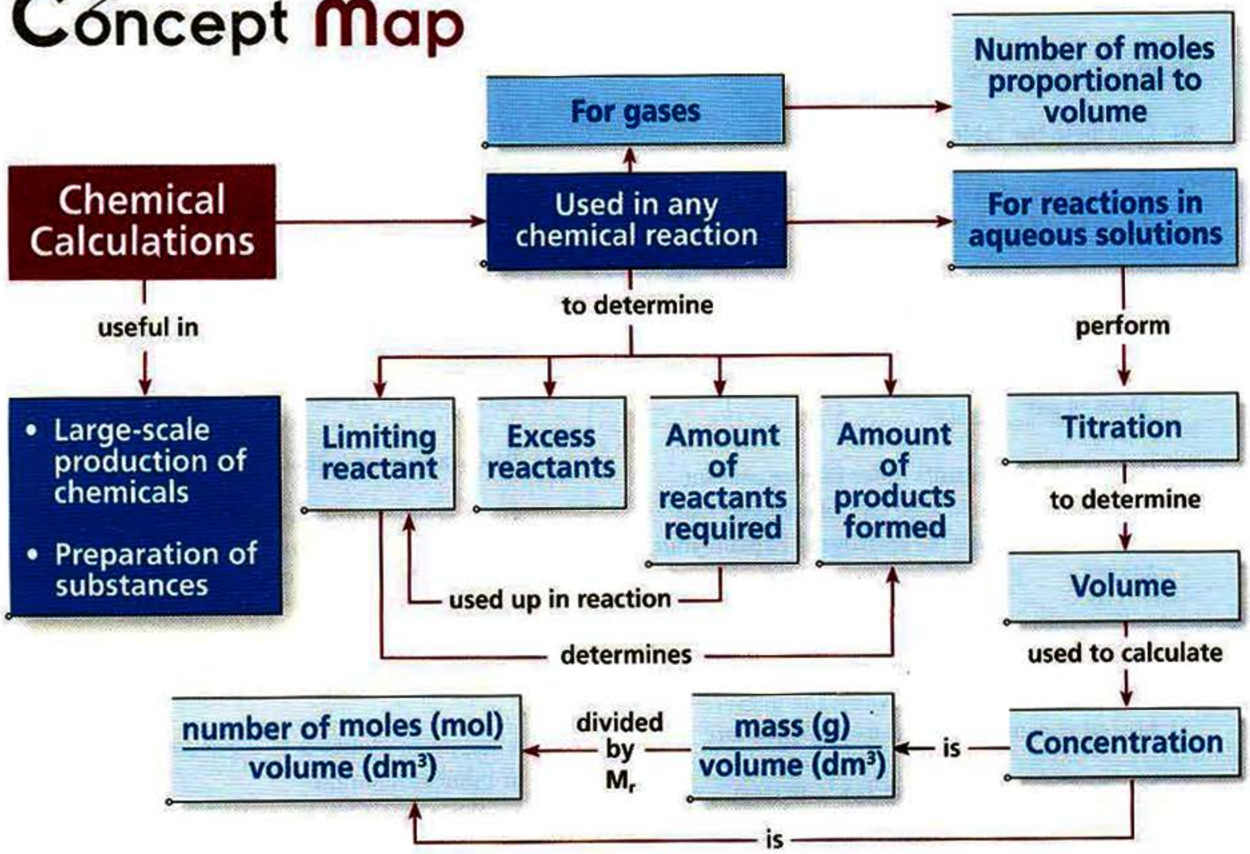
Theoretical volume of ammonia $= 0.33 \times 24 = 8.0\ dm^3$

Percentage yield of ammonia $= \dfrac{2}{8.0} \times 100\% = 25\%$

Questions

1. 3.2 g of copper was heated in air. 3.8 g of copper(II) oxide was formed. What is the percentage purity of copper?
2. 28.0 g of nitrogen reacted with 8.0 g of hydrogen to form 5.1 g of ammonia.
   a) Which reactant is in excess?
   b) What is the percentage yield of ammonia?

Concept Map



Exercise 10

Foundation

1. What is the mass of oxygen liberated when 2.13 g of sodium chlorate(V) is heated?

   $2NaClO_3(s) \longrightarrow 2NaCl(s) + 3O_2(g)$

   A 0.32 g B 0.60 g
   C 0.96 g D 1.92 g

2. When potassium hydrogencarbonate ($KHCO_3$) is heated strongly, the following reaction occurs:

   $2KHCO_3(s) \longrightarrow K_2CO_3(s) + H_2O(g) + CO_2(g)$

   What is the loss in mass when 20 g of potassium hydrogencarbonate is heated?

   A 1.8 g B 3.6 g
   C 4.4 g D 6.2 g

3. What volume of 0.1 mol/dm³ sodium hydroxide is required to exactly react with 20 cm³ of 0.1 mol/dm³ sulphuric acid? The equation for the reaction is

   $H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

   A 20 cm³ B 40 cm³
   C 60 cm³ D 80 cm³

4. Household bleach, hydrogen peroxide, decomposes to form water and oxygen.

   $2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$

   68 g of hydrogen peroxide decomposed to give 1.2 dm³ of oxygen. What was the percentage purity of the hydrogen peroxide?

   A 2.5% B 5.0%
   C 10.0% D 15.0%

5. Nitrogen reacts with hydrogen to form ammonia as shown below.

$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

a) Complete the table below to show the
 i) volume of each gas formed.
 ii) number of moles of each gas.
 iii) mass of each gas.

	N_2	H_2	NH_3
Volume	60 cm^3		
Number of moles			
Mass of gas			

b) How would the results differ for the amount of ammonia formed if 60 cm^3 of nitrogen had reacted with 90 cm^3 of hydrogen? Explain your answer.

6. a) A cook puts 4.2 g of sodium hydrogencarbonate into a cake mixture. Calculate the volume of carbon dioxide given off when the mixture is heated at room temperature and pressure.

$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$

b) Another cook makes a mixture of 9.2 g of ethanol (C_2H_5OH) in 100 cm^3 of water in which to cook some meat. What is the concentration, in mol/dm^3, of ethanol in the solution of water?

Challenge

1. Ammonia is used to produce nitric acid. The first step in the manufacturing process is the reaction between ammonia and oxygen to produce gaseous nitrogen monoxide (NO) and water vapour.

$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$

a) 1.0 mol of ammonia gas was allowed to react with 1.0 mol of oxygen. What is the limiting reactant in this reaction?

b) Using the number of moles given in (a), how many moles of each of the following will be formed?
 i) Nitrogen monoxide
 ii) Water vapour

c) Calculate the volume of nitrogen monoxide formed if 40 cm^3 of ammonia reacted with 60 cm^3 of oxygen.

2. A solution contains 2.38 g of magnesium chloride, $MgCl_2$, in 500 cm^3 of solution.

a) What is the relative formula mass of magnesium chloride?

b) Calculate the number of moles of magnesium chloride in 500 cm^3 of the solution.

c) What is the concentration of
 i) magnesium chloride in mol/dm^3?
 ii) chloride ions in g/dm^3?

3. 50 cm^3 of hydrogen peroxide, which has a concentration of 2.0 mol/dm^3, is decomposed. The equation for the reaction is

$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$

a) How many moles of hydrogen peroxide is used in the reaction?
b) How many moles of oxygen are produced in the reaction?

c) Calculate the volume of oxygen that is produced in the reaction. *(C)*

4. A solution *S* contains 15.9 g of sodium carbonate in 1 dm^3 of solution.
a) What is the concentration of *S* in mol/dm^3?
b) *S* reacts with sulphuric acid according to the reaction

$Na_2CO_3(aq) + H_2SO_4(aq)$
$\longrightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$

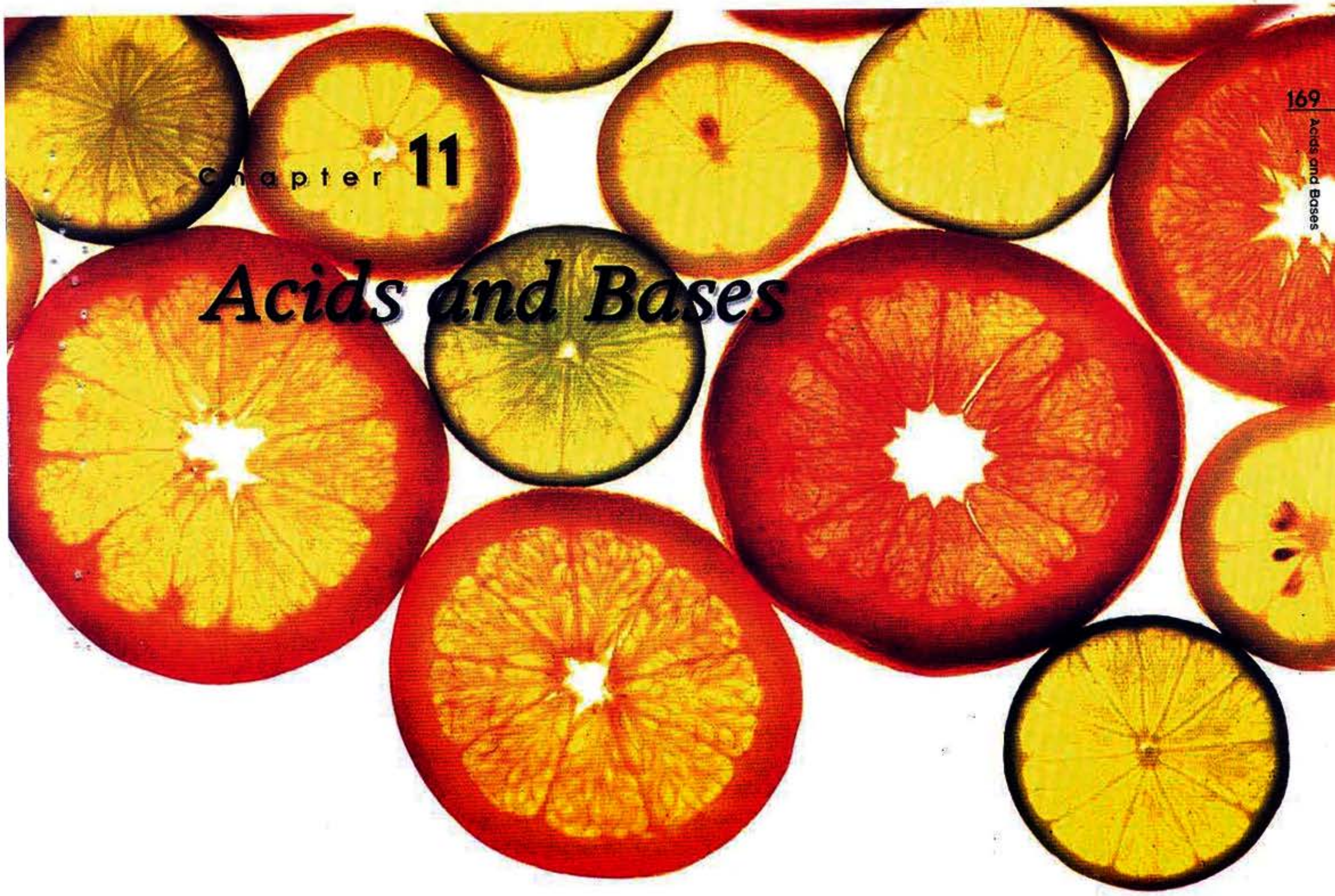
Exactly 25.0 cm^3 of *S* reacts with 12.5 cm^3 of sulphuric acid. What is the concentration of the sulphuric acid in
 i) mol/dm^3? ii) g/dm^3?

5. Sulphur dioxide dissolves in water to form sulphurous acid (H_2SO_3) as the only product.

$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$

a) Calculate the concentration of the sulphurous acid formed when 48 dm^3 of sulphur dioxide dissolved in 20 dm^3 of water.

b) Sodium hydroxide can be used to absorb sulphur dioxide. One of the products of the reaction is sodium sulphate(IV), Na_2SO_3.
 i) Write the equation for this reaction.
 ii) Calculate the volume of 0.5 mol/dm^3 sodium hydroxide required to neutralise 48 dm^3 of sulphur dioxide.



Chapter 11

Acids and Bases

Have you ever eaten lemons or oranges? If you have, then you have tasted acids! Lemons, grapefruits and oranges contain an acid called citric acid, which gives them their sour taste. In fact, the word 'acid' comes from the Latin word *acidus*, which means sour. Vinegar tastes sour because it contains ethanoic acid. Grape juice tastes sour due to tartaric acid. These are just some examples of the acids you can find in everyday items. Even your stomach produces an acid. This acid is so strong that it can burn your skin!

Your stomach is protected from the effects of this acid because it is lined by cells that produce a base. The base neutralises your stomach acid.

In this chapter, you will learn more about acids, bases and neutralisation.

Chapter Outline

11.1 | Acids

Table 11.1 below shows the names of some common acids, their formulae and the ions they produce in aqueous solution.

Name of acid	Formula	Ions produced in aqueous solution	
ethanoic acid (found in vinegar)	CH_3COOH	$H^+(aq)$	$CH_3COO^-(aq)$
hydrochloric acid	HCl	$H^+(aq)$	$Cl^-(aq)$
nitric acid	HNO_3	$H^+(aq)$	$NO_3^-(aq)$
sulphuric acid	H_2SO_4	$H^+(aq)$	$SO_4^{2-}(aq)$

Table 11.1 Common acids

Many naturally occuring acids, such as citric acid found in oranges, are sometimes known as organic acids. There are other acids called mineral acids. Most of these are man-made, e.g. nitric acid and sulphuric acid.

What is an acid?

All acids produce hydrogen ions, H^+, in aqueous solution. This leads us to the definition of **acid** as *a substance which produces hydrogen ions, H^+, when it is dissolved in water.*

All acids contain hydrogen but not all compounds that contain hydrogen are acids. For example, both ammonia, NH_3, and methane, CH_4, contain hydrogen, but they are not acids because they do not produce hydrogen ions in water. *It is the hydrogen ions produced that are responsible for the properties of acids.*

What are the properties of acids?

1. *Acids have a sour taste.*

2. *Acids dissolve in water to form solutions which conduct electricity.*

3. *Acids turn blue litmus paper red.*

4. *Acids react with reactive metals to form hydrogen and a salt* (see chapter 12). The general equation for the reaction is shown below.

   metal + acid ⟶ salt + hydrogen

Chem-Aid

A reactive metal gives up electrons easily to form positive ions. Some metals are more reactive than others, e.g. calcium is more reactive than zinc.

For example, when magnesium ribbon is added to dilute sulphuric acid, bubbles of hydrogen gas can be seen. The equation for this reaction is

magnesium + dilute sulphuric acid ⟶ magnesium sulphate + hydrogen

$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$

Magnesium sulphate is the salt produced from the above reaction. Salts are called sulphates when they are formed from sulphuric acid; nitrates when formed from nitric acid; and chlorides when formed from hydrochloric acid.

If a lighted splint is placed at the mouth of the test tube, a 'pop' sound will be heard and the lighted splint will be extinguished (Fig. 11.1). This confirms that the gas produced is hydrogen.

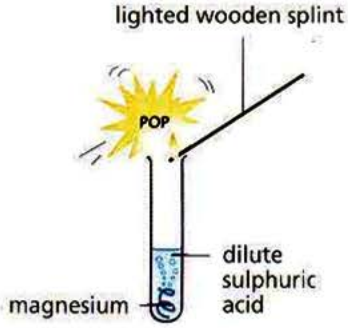


Fig. 11.1 Testing for hydrogen gas

There are some acid and metal reactions that do *not* give hydrogen.

- When unreactive metals such as copper or silver are added to dilute acids, there is no reaction.
- Concentrated nitric acid reacts with metals such as copper but it does *not* give hydrogen. Instead, a nitrate (a salt), water and nitrogen dioxide gas are formed.
- Lead appears not to react with dilute hydrochloric acid and dilute sulphuric acid. This is because a layer of lead(II) chloride or lead(II) sulphate is formed from the initial reaction between lead and the dilute acid. This layer is insoluble in water and quickly forms a coating around the metal. The coating protects the metal from further attack by the acid (Fig. 11.2).

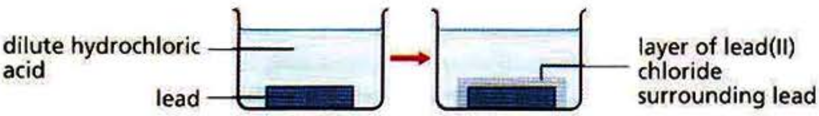


Fig. 11.2 Why lead does not seem to react with acid

Link

Why are some metals more reactive than others? Read more in chapter 14.

Quick Check

Write chemical equation for the reaction between

a) iron and dilute sulphuric acid.
b) iron and dilute hydrochloric acid.
c) magnesium and dilute hydrochloric acid.

5. *Acids react with carbonates to form a salt, carbon dioxide and water.* The general equation for the reaction is

carbonate + acid ⟶ salt + water + carbon dioxide

The reaction between sodium carbonate and dilute hydrochloric acid produces sodium chloride, water and carbon dioxide.

sodium carbonate + dilute hydrochloric acid ⟶ sodium chloride + water + carbon dioxide

$Na_2CO_3(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

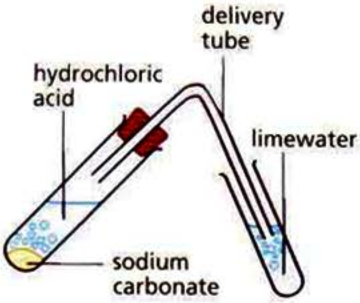


Fig. 11.3 Testing for carbon dioxide

Try it Out

Rub lemon juice onto an old copper coin. The citric acid will react with the copper(II) oxide on the coin to give you a bright shining coin!

Quick Check

Write chemical equation for the reaction between
a) copper(II) oxide and dilute sulphuric acid.
b) magnesium hydroxide and dilute hydrochloric acid.
Identify the salts in the equations you have written.

To test for carbon dioxide, bubble the gas through limewater (Fig. 11.3). Carbon dioxide forms a white precipitate with limewater.

6. *Acids react with metal oxides and hydroxides to form a salt and water only.* The general equation for the reaction of an acid with a metal oxide is

metal oxide + acid ⟶ salt + water

For example, zinc oxide reacts with dilute sulphuric acid to form zinc sulphate and water.

zinc oxide + dilute sulphuric acid ⟶ zinc sulphate + water

$$ZnO(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2O(l)$$

The general equation for the reaction of an acid with a metal hydroxide is

metal hydroxide + acid ⟶ salt + water

For example, zinc hydroxide reacts with dilute nitric acid to form zinc nitrate and water.

zinc hydroxide + dilute nitric acid ⟶ zinc nitrate + water

$$Zn(OH)_2(s) + 2HNO_3(aq) \longrightarrow Zn(NO_3)_2(aq) + 2H_2O(l)$$

All metal oxides and metal hydroxides react with acids in the same way.

What is the role of water in acidity?

If a piece of dry magnesium ribbon is added to a solution of hydrogen chloride in a dry organic solvent, there is no reaction (Fig. 11.4). When a piece of magnesium ribbon is added to a solution of hydrogen chloride in water (aqueous hydrochloric acid), bubbles of hydrogen are produced.

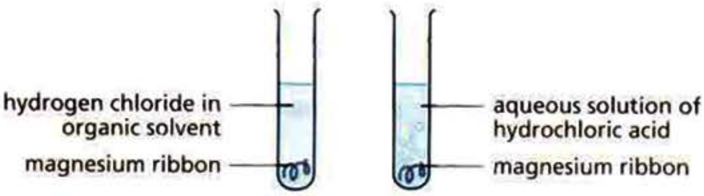


Fig. 11.4 Acids only react when they are dissolved in water.

Hydrogen chloride can exist as two different types of particles (Fig. 11.5).

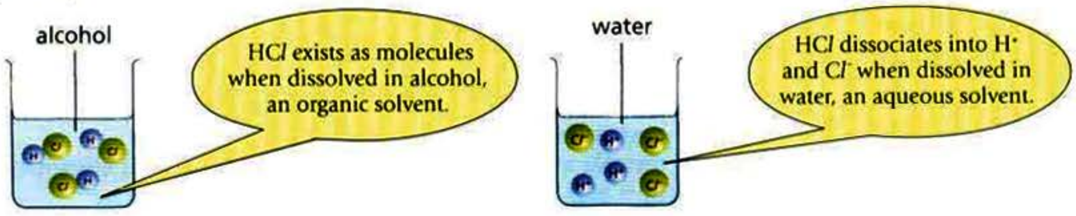


Fig. 11.5 HC*l* in organic solvent and HC*l* in water

Hydrogen chloride exists as covalent molecules. In the absence of water, for example, *in organic solvents*, they *do not behave as acids*.

Acids only show the properties of acids when they are dissolved in water. This is because acids **dissociate** in water to produce hydrogen ions which are responsible for the acidic properties.

In the case of hydrogen chloride, Fig. 11.6 shows what happens when it is dissolved in water.

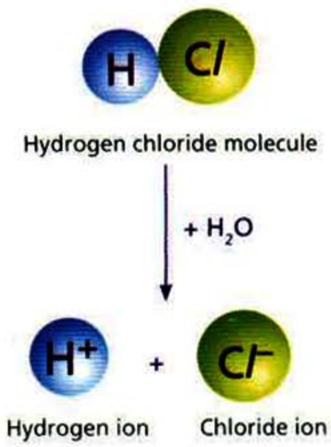


Fig. 11.6 Dissociation of hydrogen chloride in water

The hydrogen ions (H^+) produced allow acids to react with metals such as magnesium, as shown by the equations:

hydrogen ions + magnesium ⟶ hydrogen gas + magnesium ions

$2H^+(aq) + Mg(s) \longrightarrow H_2(g) + Mg^{2+}(aq)$

No dissociation occurs when hydrogen chloride is dissolved in an organic solvent (Fig.11.5). Thus, hydrogen chloride dissolved in an organic solvent does not react with magnesium.

Similarly, no reaction occurs when solid citric acid is mixed with anhydrous sodium carbonate. However, when a few drops of water are added to the mixture, bubbles of carbon dioxide are produced.

hydrogen ions + carbonate ions ⟶ water + carbon dioxide gas

$2H^+(aq) + CO_3^{2-}(aq) \longrightarrow H_2O(l) + CO_2(g)$

Chem-Aid

1. When a substance **dissociates**, it dissolves in water to give a solution that contains ions.
2. An **anhydrous** compound is one with all water of crystallisation removed.

Uses of Acids

The uses of some acids are shown in Fig. 11.7.

Uses of acids

Sulphuric acid:
- In making detergents
- In making fertilisers
- In car batteries

Hydrochloric acid:
- In leather processing
- For cleaning metals

Ethanoic acid:
- (In vinegar) to preserve food
- In making adhesives such as glue

Fig. 11.7 Uses of some acids

Key ideas

1. An acid is a substance which produces hydrogen ions (H^+) when dissolved in water.
2. The main properties of acids are as follows. They
   - have a sour taste,
   - dissolve in water to form solutions which conduct electricity,
   - turn blue litmus paper red,
   - react with reactive metals to produce hydrogen gas and a salt,
   - react with carbonates to produce carbon dioxide gas, water, and a salt,
   - react with metal oxides and hydroxides to produce a salt and water only.

Test Yourself 11.1

Worked Example

A liquid *T* has a sour taste and it changed the colour of blue litmus paper to red. Give two different chemical tests, and the results of each test, to confirm that *T* is an acid.

Thought Process

Recall the properties of an acid and the products of its reactions:

- An acid reacts with a reactive metal to form a salt and hydrogen gas.
- An acid reacts with a carbonate to produce a salt, water and carbon dioxide gas.

Answer

For the first test, add magnesium ribbon to liquid *T* and test for hydrogen gas. The lighted splint is extinguished with a 'pop' sound.

For the second test, add calcium carbonate to liquid *T* and test for carbon dioxide gas. The gas produces a white precipitate with limewater.

Questions

1. Suggest why barium reacts with dilute hydrochloric acid but appears **not** to react with dilute sulphuric acid.
2. When solid *X* is added to solution *Y*, copper(II) chloride and carbon dioxide gas are formed. However, when solid *Z* is added to solution *Y*, zinc chloride and hydrogen are produced. Deduce the identities of *X*, *Y* and *Z*.
3. Explain why solid citric acid does not conduct electricity but a solution of citric acid in water does.
4. Write the ionic equation for the reaction between
   a) zinc and an acid.
   b) calcium carbonate and an acid.

11.2 | Bases and Alkalis

Bases

What is a base?

A base is *any metal oxide or hydroxide*. This means that a base contains either oxide ions, O^{2-}, or hydroxide ions, OH^-.

Table 11.2 gives the names and formulae of some common bases.

Name of base	Formula
sodium oxide	Na_2O
zinc oxide	ZnO
copper(II) oxide	CuO
magnesium hydroxide	$Mg(OH)_2$
aluminium hydroxide	$Al(OH)_3$

Table 11.2 Some common bases

We can also define a **base** as *a substance that reacts with an acid to give a salt and water only.* The general equation for this reaction is

base + acid ⟶ salt + water

For example,

copper(II) oxide + dilute sulphuric acid ⟶ copper(II) sulphate + water

$$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$$

The ionic equation for this reaction is

$$CuO(s) + 2H^+(aq) \longrightarrow Cu^{2+}(aq) + H_2O(l)$$

Consider the reaction between sodium hydroxide and dilute nitric acid.

sodium hydroxide + dilute nitric acid ⟶ sodium nitrate + water

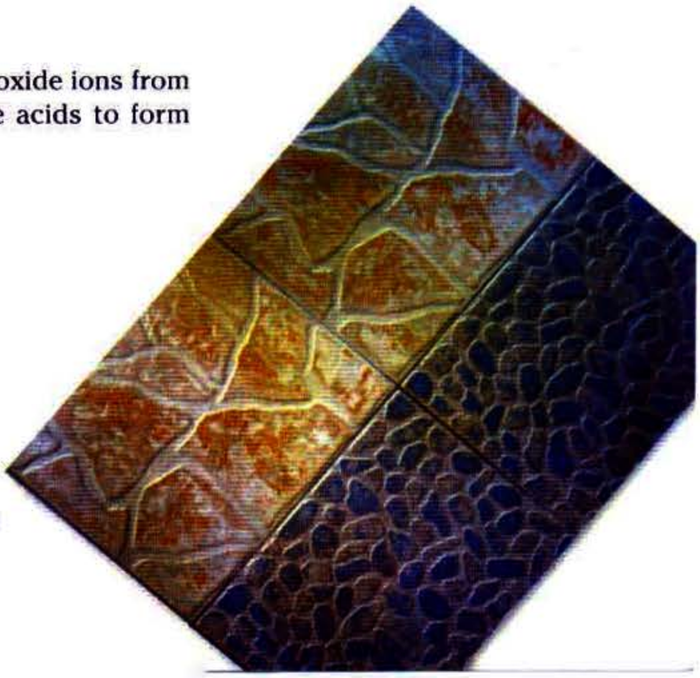
$$NaOH(aq) + HNO_3(aq) \longrightarrow NaNO_3(aq) + H_2O(l)$$

The ionic equation for this reaction is

$$OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$$

In the above reactions, the oxide ions or the hydroxide ions from the bases react with the hydrogen ions from the acids to form water.

These are antacids. They are used to ease stomachaches due to excess acid in the stomach. What type of compound is used to make antacids?



Aluminium hydroxide is used to manufacture ceramics.

Alkalis: A Special Class of Bases

An **alkali** is *a base that is soluble in water*. An example of a soluble base is sodium oxide.

sodium oxide + water ⟶ sodium hydroxide

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

Most bases are insoluble in water. Copper(II) oxide does not react nor dissolve in water. Thus, it is a base and not an alkali. Table 11.3 shows some common alkalis.

Chem-Aid

1. All alkalis are bases but not all bases are alkalis.
2. Sodium hydroxide is sometimes called caustic soda. Caustic means that the chemical substance burns or destroys living tissue.

Name of alkali	Formula
sodium hydroxide	$NaOH$
potassium hydroxide	KOH
calcium hydroxide	$Ca(OH)_2$
barium hydroxide	$Ba(OH)_2$
aqueous ammonia	NH_3

Table 11.3 Some common alkalis

What are the properties of alkalis?

1. *Alkalis have a bitter taste and soapy feel.*
2. *Alkalis turn red litmus paper blue.*
3. *All alkalis produce hydroxide ions when dissolved in water.*

   For example,

   sodium hydroxide ⟶ sodium ion + hydroxide ion

   $$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

   calcium hydroxide ⟶ calcium ion + hydroxide ion

   $$Ca(OH)_2(aq) \longrightarrow Ca^{2+}(aq) + 2OH^-(aq)$$

   When ammonia gas dissolves in water, ammonium ions and hydroxide ions are formed.

   ammonia + water ⇌ ammonium ion + hydroxide ion

   $$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

4. *All alkalis can react with acids to form a salt and water only.* This reaction is called **neutralisation**.

   In a neutralisation reaction, the hydrogen ions from the acid and the hydroxide ions from the alkali react to form water.

   For example,

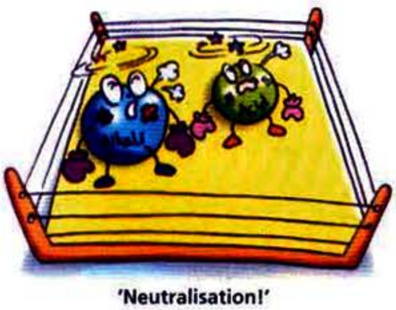
   sodium hydroxide + hydrochloric acid ⟶ sodium chloride + water

   $$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

   The ionic equation for this reaction is

   $$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

   This is the ionic equation for any neutralisation reaction.



'Neutralisation!'

Consider the reaction of sodium hydroxide and sulphuric acid.

sodium hydroxide + sulphuric acid ⟶ sodium sulphate + water

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

The ionic equation for this reaction is also

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Now, consider the following reactions which involve reacting manganese(II) oxide, MnO and manganese(IV) oxide, MnO_2, with hydrochloric acid, HC*l*.

manganese(II) oxide + hydrochloric acid ⟶ manganese(II) chloride + water

$$MnO(s) + 2HCl(aq) \longrightarrow MnCl_2(aq) + H_2O(l) \quad \text{-------------------- (1)}$$

manganese(IV) oxide + hydrochloric acid ⟶ manganese(II) chloride + water + chlorine

$$MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g) \quad \text{---------- (2)}$$

Which one of the above reactions is NOT a neutralisation reaction?

Reaction (2) is *not* a neutralisation reaction. In this reaction, besides the salt, manganese(II) chloride, and water, chlorine is also produced. Reaction (1) is a neutralisation reaction because *only* the salt, manganese(II) chloride, and water are produced.

5. *Alkalis heated with ammonium salts give off ammonia gas*. The general equation for this reaction is

alkali + ammonium salt ⟶ ammonia + water + salt

Ammonia can be recognised by its characteristic pungent smell. It also turns moist red litmus paper blue (Fig. 11.8).

For example, when calcium hydroxide is heated with ammonium chloride, ammonia gas is produced.

calcium hydroxide + ammonium chloride ⟶ calcium chloride + water + ammonia

$$Ca(OH)_2(aq) + 2NH_4Cl(s) \longrightarrow CaCl_2(aq) + 2H_2O(l) + 2NH_3(g)$$

When alkalis react with ammonium salts, the hydroxide ions and the ammonium ions react to produce ammonia gas. The ionic equation for the reaction is

hydroxide ion + ammonium ion ⟶ ammonia + water

$$OH^-(aq) + NH_4^+(aq) \longrightarrow NH_3(g) + H_2O(l)$$

Quick Check

Name the reactants used to prepare ammonium sulphate in the laboratory. Write the chemical equation for the reaction.

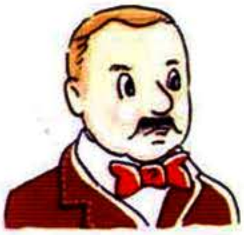
Try it Out

Find out which (alkali) ingredient in toothpaste helps prevent tooth decay. Write the equation for the neutralisation reaction that occurs.

Chem-Aid

Manganese has variable oxidation states. It can form more than one oxide and each oxide has different reactions with acids. You will learn more about oxidation state in chapter 13.

Fig. 11.8 Testing for ammonia



Svante Arrhenius (1859 – 1927)

Arrhenius is best known for the theory that led to the definition of acids and bases. He also proposed the theory of neutralisation. At first, his ideas were not popular with the well-known chemists of his time. He had to produce a lot of experimental data and logical arguments before he could finally convince them. Arrhenius had the ability to make science interesting to the general public. His lectures were very well attended. Svante Arrhenius received the Nobel Prize for Chemistry in 1903.

Link

Which other cations can be identified using alkalis? Find out in chapter 12.

Did you know that cleaning agents such as soap and many household detergents contain alkalis?

6. *Alkalis can react with a solution of one metal salt to give metal hydroxide and another metal salt.* The general equation for this reaction is

alkali + salt (of metal *A*) ⟶ metal hydroxide + salt (of metal *B*)

The metal hydroxide appears as a precipitate if it is insoluble in water.

For example, sodium hydroxide reacts with a solution of iron(II) sulphate to give a green precipitate of iron(II) hydroxide.

sodium hydroxide + iron(II) sulphate ⟶ iron(II) hydroxide + sodium sulphate

$$2NaOH(aq) + FeSO_4(aq) \longrightarrow Fe(OH)_2(s) + Na_2SO_4(aq)$$

In the laboratory, alkalis such as aqueous sodium hydroxide and aqueous ammonia can be used to test for cations (e.g. Fe^{2+}, Al^{3+}, etc.) in metal salts.

Uses of Bases and Alkalis

The uses of some bases and alkalis are shown in Fig 11.9.

Uses of bases and alkalis

Ammonia solution:
- In window cleaning solutions
- In fertilisers

Calcium oxide:
- In neutralising acidic soil
- To make iron, concrete and cement

Magnesium hydroxide:
- In toothpaste to neutralise acid on teeth
- In antacids, to relieve indigestion

Sodium hydroxide:
- In making soaps and detergents
- In industrial-cleaning detergents

Fig. 11.9 Uses of bases and alkalis

Key ideas

1. A base is a substance which reacts with acids to form a salt and water only.
2. An alkali is the solution formed when a base dissolves in water.
3. Alkalis produce hydroxide ions (OH^-) when dissolved in water.
4. The main properties of alkalis are as follows. They
   - have a bitter taste and a soapy feel,
   - turn red litmus paper blue,
   - produce ammonia gas when heated with ammonium salts,
   - react with a solution of one metal salt to give metal hydroxide and another metal salt.
5. When an acid is neutralised by an alkali, the ionic equation is

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Test Yourself 11.2

Worked Example

Which two substances react to give a salt and water only?

A Copper(II) oxide and ethanoic acid.
B Magnesium and sulphuric acid.
C Sodium oxide and water.
D Zinc carbonate and hydrochloric acid.

Thought Process

A: Involves a metal oxide and an acid. A base (metal oxide or hydroxide) reacts with an acid to produce a salt and water. This is called neutralisation.
B: Involves a reactive metal and an acid. A reactive metal reacts with an acid to produce a salt and hydrogen gas.
C: Involves a base and water. A soluble base dissolves in water to form an alkaline solution.
D: Involves a carbonate and an acid. A carbonate reacts with an acid to form a salt, water and carbon dioxide.

Answer

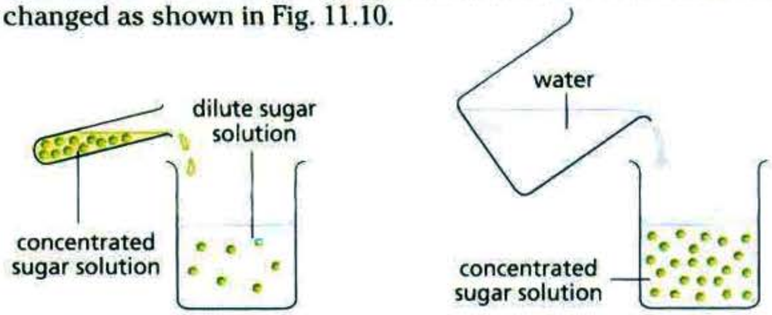
A

Questions

1. Name **two** bases that react with dilute sulphuric acid to give zinc sulphate.
2. State a main difference between acids and alkalis.
3. Stomach acid contains dilute hydrochloric acid. Write a chemical equation to show the neutralisation of stomach acid by milk of magnesia, $Mg(OH)_2$.

11.3 | Concentration and Strength

The term **concentration** tells us *how much a substance is dissolved in 1 dm^3 of the solution.* The concentration of a solution can be changed as shown in Fig. 11.10.



(a) Increasing concentration of dilute sugar solution by adding concentrated sugar solution

(b) Decreasing concentration of concentrated sugar solution by adding water

Fig. 11.10 Changing the concentration of a solution

Similarly, the concentration of an acid or an alkali can be changed. Unlike their concentration, the strength of an acid or an alkali cannnot be changed.

The term **strength** refers to *how easily an acid or an alkali dissociates when dissolved in water.* A strong acid, like hydrochloric acid, dissociates easily in water. A weak acid, like ethanoic acid, does not fully dissociate when dissolved in water. It has a much lower concentration of hydrogen ions in solution. Since the strength of an acid or alkali cannot be changed, hydrochloric acid is a strong acid whether it is diluted or concentrated. Ethanoic acid remains a weak acid, whether it is diluted or concentrated.

The strength of an acid or alkali can be shown using the pH scale.

11.4 | The pH Scale

What is the pH scale?

The pH scale is a set of numbers used to indicate whether a solution is acidic, neutral or alkaline. For example, acids have pH values less than 7; alkalis have pH values greater than 7; and a neutral solution has a pH value of exactly 7. The pH of some common substances is shown in Fig. 11.11.

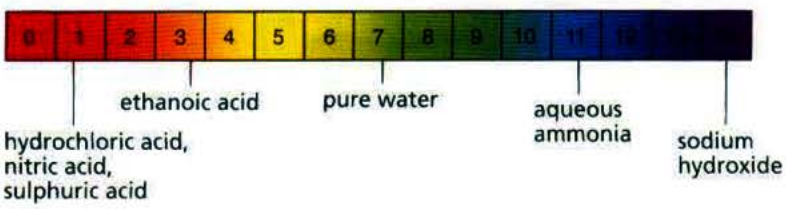


Fig. 11.11 pH values of common substances

How can we determine the pH of a given solution?
The pH of a given solution can be determined by using
1. a chemical compound called an indicator (Fig. 11.12),
2. a pH probe attached to a data logger (Fig. 11.13).

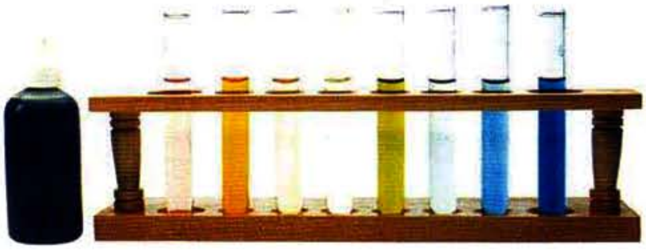


Fig. 11.12 The colour of Universal Indicator (paper or solution) is different in solutions of different pH.

Fig. 11.13 Using a pH probe and a data logger to measure pH

Universal Indicator

The pH value of a solution can be determined by using Universal Indicator. Universal Indicator contains a mixture of dyes. It gives different colours in solutions of different pH (Fig. 11.14).

How is the pH of a solution related to the concentration of hydrogen ions in the solution?
The pH of a solution is calculated based on the number of hydrogen ions or hydroxide ions present in a solution. Acids with a smaller pH value have a higher concentration of hydrogen ions. Alkaline solutions with a larger pH value have a higher concentration of hydroxide ions.

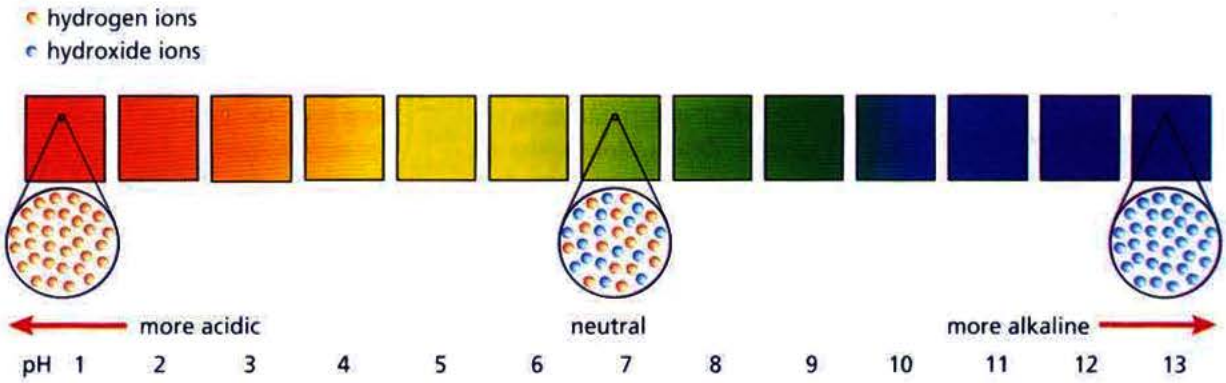


Fig. 11.14 The pH of a solution is related to the concentration of hydrogen ions and hydroxide ions.

TidBit

Many foods contain a natural dye called turmeric. Turmeric is a natural indicator. When added to vinegar (which contains ethanoic acid), turmeric is orangey-yellow. Turmeric stains cannot be easily washed off your clothes. Washing powder dissolved in water forms an alkaline solution that turns the turmeric stains pink instead of removing them!

Hence, pH can be used to compare the strength of acids and alkalis of the *same* concentration. For example, 0.1 mol/dm^3 of hydrochloric acid has a pH of 1. It is a stronger acid than 0.1 mol/dm^3 of ethanoic acid, which has a pH of 3.

Besides Universal Indicator, there are many other coloured indicators that are commonly used in titrations. Each indicator has one colour in an acidic solution and another colour in an alkaline solution. It will change its colour at a certain pH value. Some important indicators are listed in Table 11.4.

Indicator	Colour in acidic solution	pH range at which indicator changes colour	Colour in alkaline solution
methyl orange	red	3 – 5	yellow
screened methyl orange	violet	3 – 5	green
litmus	red	5 – 8	blue
bromothymol blue	yellow	6 – 8	blue
phenolphthalein	colourless	8 – 10	pink

Table 11.4 Colour changes of some indicators

The Importance of pH

Your blood is slightly alkaline as it has a pH of about 7.4. When a person is given an injection, the substance being injected must have a pH of almost 7.4. If the pH of the blood is changed by one unit, say to 8.3, the person would die! In what other ways is pH important?

Why is soil pH important?

The pH of soil can vary from 4 to 8, depending on the type of soil. The pH of soil in Singapore's mangrove forests is neutral to slightly acidic. This is due to the presence of acidic clay.

It is important to control the pH of soil because this affects the growth and development of plants. Most plants grow best when the soil is neutral or slightly acidic (around pH 6 or 7). Plants will not grow in soil that is too acidic. This can happen when too much fertiliser is added to the soil. Another reason why soil becomes acidic is due to acid rain.

Table 11.5 shows the optimum pH range for some crops. Nothing will grow if the pH is below 5 or above 9.

Crop	Optimum pH range
potatoes, peanuts	5.5 – 6.5
wheat, plums	6.5 – 7.5
barley, cabbages	7.5 – 8.5

Table 11.5 The optimum pH range for healthy growth of some crops

How would we control the acidity of soil?

Chemicals are often added to the soil to adjust its pH. In areas where the soil is too acidic, it can be treated with bases such as quicklime (calcium oxide) or slaked lime (calcium hydroxide). This is known as 'liming' the soil. The bases react with the acids in the soil and raise the pH so that plants can grow healthily. However, adding too much base will make the soil too alkaline and unsuitable for plant growth.

Science Skills

Using simple laboratory apparatus, how would you try to determine the pH of a sample of soil?

Key ideas

1. The pH scale is used to determine if a given substance is acidic, alkaline or neutral.
2. The pH of colourless solutions can be determined using Universal Indicator.
3. It is important to control the pH of soil. Most plants grow best when the pH is around 7.
4. Quicklime (calcium oxide) and slaked lime (calcium hydroxide) are commonly used to reduce the acidity of soil.

Link

How does acid rain affect the environment? Read more in chapter 20.

Test Yourself 11.3

Questions

1. A farmer had a poor growth of barley last year because his soil was too acidic. Suggest what he might do to improve the growth of barley this year.

2. An iron kettle is coated with 'scales' of calcium carbonate.
   a) Explain why you can use lemon juice for 'descaling'.
   b) Name an acid commonly used in cooking that can be used to 'descale' the kettle.

3. The table below gives the pH of various aqueous solutions.

Solution	*U*	*V*	*W*	*X*	*Y*	*Z*
pH	2	1	9	5	6	11

   a) Give **two** pairs of solutions that would react together to give a solution with a pH of 7.
   b) Three of the substances *U, V, W, X, Y,* and *Z* are aqueous ammonia, nitric acid and carbonic acid. Which letters correspond to these three substances?

11.5 | Types of Oxides

Many acids and alkalis are formed by dissolving oxides in water. An oxide is a compound of oxygen and another element. Most oxides can be grouped into four types: acidic oxides, basic oxides, amphoteric oxides and neutral oxides.

Acidic Oxides

Non-metals may form **acidic oxides**. Most acidic oxides dissolve in water to form an acid. One example is sulphur dioxide, which dissolves readily in water to form sulphurous acid.

sulphur dioxide + water ⟶ sulphurous acid

$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$

Carbonated drinks are slightly acidic due to the carbon dioxide dissolved in them. The carbon dioxide is added to give 'fizz' to the drinks, i.e. the gas bubbles that readily move through and escape from the drinks.

Other examples of acidic oxides are given in Table 11.6.

Acidic oxide	Formula	Physical state	Acid produced in water	
			Name	Formula
carbon dioxide	CO_2	gas	carbonic acid	H_2CO_3
sulphur trioxide	SO_3	gas	sulphuric acid	H_2SO_4
phosphorus(V) oxide	P_4O_{10}	solid	phosphoric acid	H_3PO_4

Table 11.6 Examples of acidic oxides

Acidic oxides do not react with acids. However, they react with alkalis to form a salt and water. For example

sulphur dioxide + sodium hydroxide ⟶ sodium sulphite + water

$SO_2(g) + 2NaOH(aq) \longrightarrow Na_2SO_3(aq) + H_2O(l)$

Is silicon(IV) oxide an acidic oxide?
Silicon(IV) oxide, SiO_2, is a solid at room temperature. It does not dissolve in water, but it reacts with sodium hydroxide to form sodium silicate (a salt).

silicon(IV) oxide + sodium hydroxide ⟶ sodium silicate + water

$SiO_2(s) + 2NaOH(aq) \longrightarrow Na_2SiO_3(aq) + H_2O(l)$

Hence, silicon(IV) oxide is an acidic oxide.

Basic Oxides

The *oxides of metals are* **basic oxides**. Most basic oxides are insoluble in water. A few oxides, such as sodium oxide and potassium oxide, dissolve readily in water. They are called **alkalis**.

When calcium oxide (quicklime) is added to water, a vigorous reaction occurs. Calcium hydroxide is formed, which is sparingly soluble in water. The solution of calcium hydroxide in water is called limewater.

calcium oxide + water ⟶ calcium hydroxide

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$$

Basic oxides are solids at room temperature. *They react with acids to form a salt and water.* For example:

calcium oxide + nitric acid ⟶ calcium nitrate + water

$$CaO(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + H_2O(l)$$

Amphoteric Oxides

Look at the reaction between hydrochloric acid and zinc oxide, and the reaction between sodium hydroxide solution and zinc oxide.

hydrochloric acid + zinc oxide ⟶ zinc chloride + water

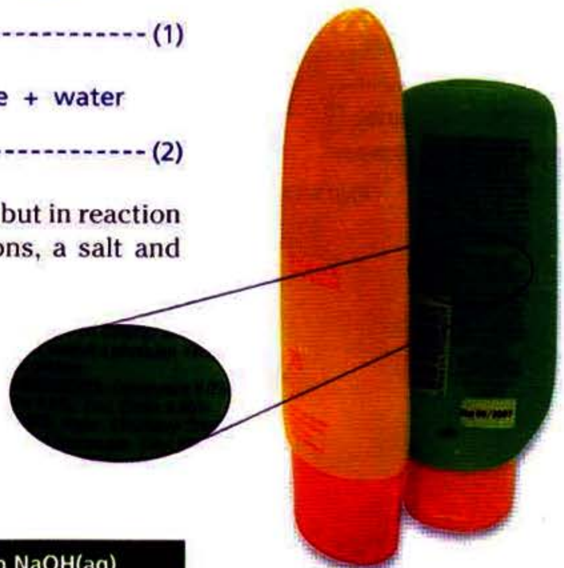
$$2HCl(aq) + ZnO(s) \longrightarrow ZnCl_2(aq) + H_2O(l) \text{ ----------------- (1)}$$

sodium hydroxide + zinc oxide ⟶ sodium zincate + water

$$2NaOH(aq) + ZnO(s) \longrightarrow Na_2ZnO_2(aq) + H_2O(l) \text{ ------------- (2)}$$

In reaction (1), zinc oxide behaves as a basic oxide, but in reaction (2), it behaves as an acidic oxide! In both reactions, a salt and water are formed.

Zinc oxide is an example of an **amphoteric oxide**. Amphoteric oxides *are metallic oxides that react with both acids and bases to form salts and water.* Other examples of amphoteric oxides are given in Table 11.7.



Zinc oxide is used in some lotions and creams to protect against sunburn as it absorbs ultraviolet light.

Amphoteric oxide	Formula	Salt produced in NaOH(aq)
aluminium oxide	Al_2O_3	sodium aluminate ($NaAlO_2$)
lead(II) oxide	PbO	sodium plumbate(II) (Na_2PbO_2)

Table 11.7 Examples of amphoteric oxides

Neutral oxide	Formula
water	H_2O
carbon monoxide	CO
nitric oxide	NO

Table 11.8 Examples of neutral oxides

Neutral Oxides

Some non-metals form oxides that show neither basic nor acidic properties. These oxides are called **neutral oxides** and they are insoluble in water. Some examples of neutral oxides are given in Table 11.8.

How can we classify an unknown oxide?

If you were given a sample containing an oxide, how would you determine if it is basic, acidic or amphoteric? Below is a step-by-step guide on how you can determine the nature of the oxide.

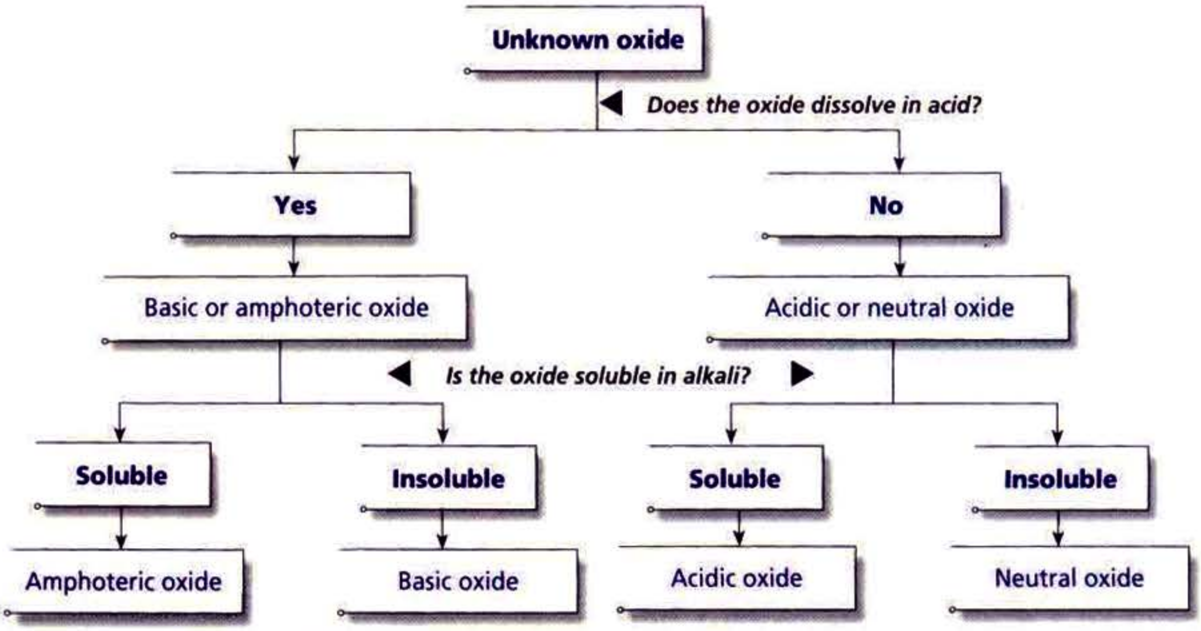


Fig. 11.15 Step-by-step guide for determining the nature of an unknown oxide

Key ideas

1. Metals react with oxygen to form basic oxides or amphoteric oxides.
2. Non-metals react with oxygen to form acidic oxides or neutral oxides.
3. Basic oxides that are soluble in water are called alkalis.
4. Basic oxides that are insoluble in water are called insoluble bases.

Test Yourself 11.4

Questions

1. Factory chimneys are often lined with calcium oxide to prevent waste gases from polluting air. Name **three** waste gases that could be removed by calcium oxide.

2. Using the observations and the flow chart on page 186, classify the oxides *A*, *B* and *C* as either amphoteric, basic or acidic. Suggest possible identities for *A*, *B* and *C*.
   A: A white solid, insoluble in water, forms a colourless solution with both $HCl(aq)$ and $NaOH(aq)$
   B: A white solid, soluble in water, forms a solution which turns blue litmus paper red
   C: A white solid dissolves sparingly in water to form a white suspension. The solution turns red litmus paper blue

11.6 | Sulphur Dioxide and Sulphuric Acid

Sulphur dioxide, an acidic oxide, has many uses, of which the most important is the manufacture of sulphuric acid. In this section, we take a closer look at how the properties of sulphur dioxide and sulphuric acid determine their uses.

TidBit

The first step in the manufacture of sulphuric acid is burning sulphur in oxygen to form sulphur dioxide, SO_2. Sulphur dioxide is further reacted with oxygen to form sulphur trioxide, SO_3.

Sulphur trioxide is first dissolved in concentrated sulphuric acid to give a fuming liquid called oleum, $H_2S_2O_7$. Water is then added to oleum to form concentrated sulphuric acid.

The two-step equations for the conversion of sulphur trioxide to sulphuric acid are $SO_3(g) + H_2SO_4(l) \longrightarrow H_2S_2O_7(l)$ and $H_2S_2O_7(l) + H_2O(l) \longrightarrow 2H_2SO_4(l)$

Properties and Uses of Sulphur Dioxide

1. **As a bleaching agent**
   A bleaching agent decolourises coloured compounds, causing them to turn pale or white.

   Sulphur dioxide bleaches coloured compounds by removing oxygen from them. Substances such as sulphur dioxide that remove oxygen from other substances are known as reducing agents (which we will learn more about in chapter 13).

The wood pulp used for making paper is coloured because it contains dyes. These dyes contain oxygen. When sulphur dioxide is added to the pulp, it removes oxygen from the dyes, causing the pulp to turn pale or white.

2. **As a food preservative**
   Sulphur dioxide is poisonous to all organisms, especially bacteria. In the food industry, sulphur dioxide is added to food in small amounts to prevent the growth of moulds and bacteria. The minute concentrations of sulphur dioxide in foodstuff are sufficient to kill bacteria but not humans. Since sulphur dioxide is poisonous, food manufacturers need to strictly control the amounts added.

Properties and Uses of Sulphuric Acid

Sulphuric acid is one of the most widely used acids. It is, after all, the cheapest and most readily available strong acid.

Link

What effects does sulphur dioxide in the air have on health and the environment? Find out in chapter 20.

1. **Manufacture of fertilisers**
   The most important use of sulphuric acid is in manufacturing fertilisers such as ammonium sulphate, $(NH_4)_2SO_4$, and superphosphate.

   Ammonium sulphate (a salt) is formed when sulphuric acid is reacted with the alkali, ammonia.

   ammonia + sulphuric acid ⟶ ammonium sulphate + water

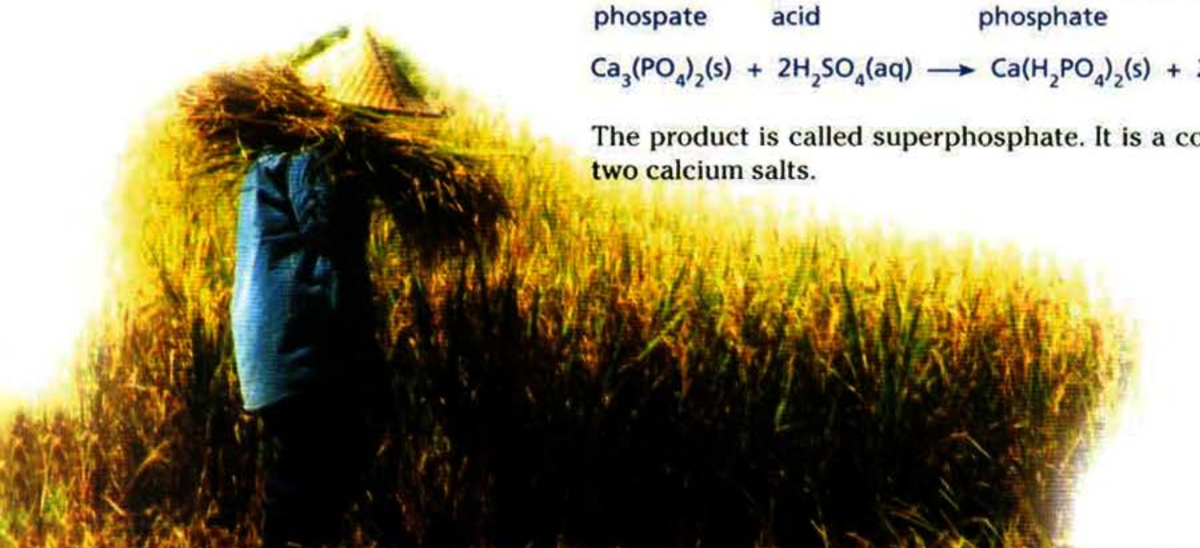
   $$2NH_3(aq) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq)$$

   Superphosphate is manufactured by reacting the raw material, calcium phosphate $Ca_3(PO_4)_2$, with concentrated sulphuric acid.

   calcium phospate + sulphuric acid ⟶ calcium dihydrogen phosphate + calcium sulphate

   $$Ca_3(PO_4)_2(s) + 2H_2SO_4(aq) \longrightarrow Ca(H_2PO_4)_2(s) + 2CaSO_4(s)$$

   The product is called superphosphate. It is a combination of two calcium salts.



Crops need phosphorus and nitrogen to grow well — ammonium sulphate provides nitrogen while superphosphate provides phosphorus.

2. **Manufacture of detergents**
Concentrated sulphuric acid is also used to manufacture detergents. It is used to treat compounds called hydrocarbons to first form an organic acid.

hydrocarbon + concentrated sulphuric acid ⟶ organic acid

The organic acid is then neutralised with sodium hydroxide solution to produce the detergent.

organic acid + sodium hydroxide ⟶ detergent

3. **As battery acid in cars**
Dilute sulphuric acid is used in batteries for cars. Lead plates and lead(IV) oxide plates are also fitted in the batteries. When sulphuric acid, lead and lead(IV) oxide react, electrical energy is produced. The electrical energy starts the car engine running.

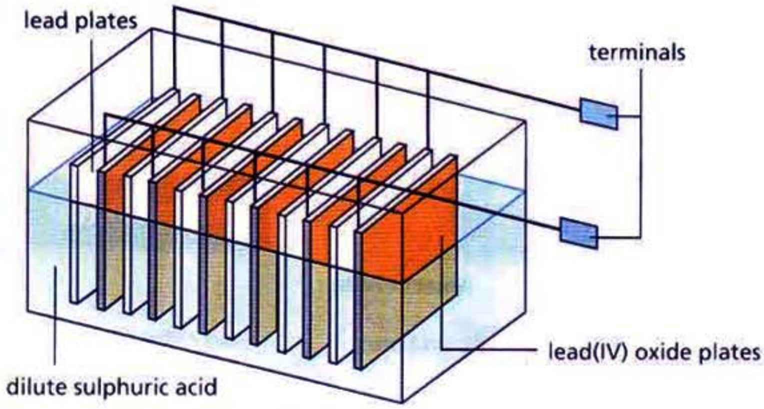


Fig. 11.16 The structure of a car battery

Sulphuric acid is also used for making synthetic fibres and for making paints and pigments. Fig. 11.17 summarises the main uses of sulphuric acid.

Key ideas

1. Sulphur dioxide is used
   - for manufacturing sulphuric acid,
   - as a bleaching agent,
   - as a food preservative.
2. Sulphuric acid is used
   - for manufacturing fertilisers,
   - for manufacturing detergents,
   - in car batteries.

Link

Hydrocarbons are compounds made up of the elements hydrogen and carbon only. They are usually extracted from crude oil. What is crude oil and what other important uses do hydrocarbons have? Find out in chapter 21.

A car battery

Fig. 11.17 The main uses of sulphuric acid

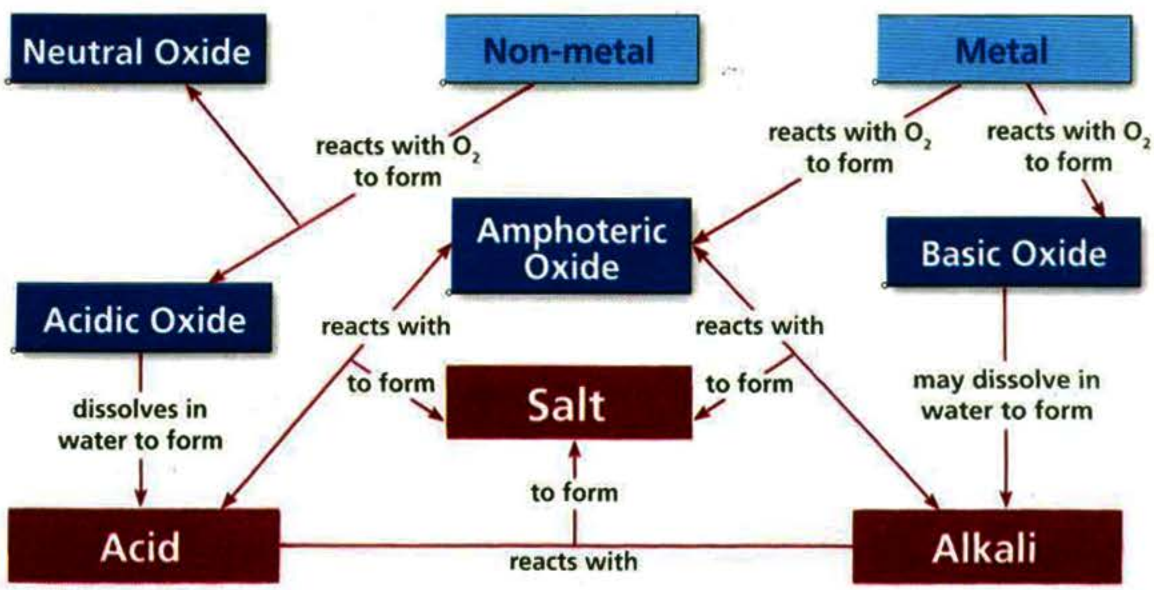
Test Yourself 11.5

Questions

1. Which of the following is not a use of sulphur dioxide?
   A Sterilising water. B Making bleaches.
   C Manufacturing sulphuric acid. D Preserving fruit drinks.

2. Sulphuric acid is not used in the manufacture of ________ .
   A detergent B ammonium sulphate
   C soap D superphosphate

Concept Map



Exercise 11

Foundation

1. A piece of magnesium ribbon does **not** react when put into a solution of hydrogen chloride in methylbenzene, an organic solvent. Which change will cause this reaction to occur?
   - **A** Adding more methylbenzene.
   - **B** Adding water and stirring.
   - **C** Stirring the mixture vigorously.
   - **D** Warming the mixture.

2. Which statement is true for all acids in aqueous solution?
   - **A** They have pH values greater than 7.
   - **B** They react with copper to liberate hydrogen.
   - **C** They react with sodium carbonate to liberate carbon dioxide.
   - **D** They turn moist red litmus paper blue.

3. Which compound neutralises the acidity of soil?
   - **A** Ammonium chloride
   - **B** Calcium hydroxide
   - **C** Magnesium sulphate
   - **D** Potassium nitrate

4. Aluminium oxide is amphoteric. Therefore, solid aluminium oxide ________.
   - **A** is reduced by carbon
   - **B** is soluble in water
   - **C** reacts with sodium hydroxide to form a salt
   - **D** will conduct electricity

5. To remove an oil stain from a shirt, baking powder was first rubbed onto the stain. The same oil stain was then rubbed with a detergent solution and a pungent smell of ammonia was produced. What substances were likely present in the baking powder and detergent respectively to cause this?
   - **A** Sodium chloride and sodium hydroxide.
   - **B** Ammonium carbonate and sodium chloride.
   - **C** Ammonium carbonate and sodium hydroxide.
   - **D** Calcium carbonate and magnesium sulphate.

6. a) Explain, with an example in each case, what is meant by
      i) an alkali.
      ii) an acid.
      iii) neutralisation.

   b) Boron oxide, B_2O_3, is an acidic oxide. State **two** inferences you can make from this statement.

7. State, with reasons, whether each of the following statements is **true** or **false**.
   a) All alkalis are bases.
   b) All acids contain oxygen.
   c) All compounds that contain hydrogen have a pH less than 7.
   d) Aqueous solutions of acids conduct electricity.

8. Which of the substances in the list below have the uses stated?

calcium hydroxide	sulphur dioxide	nitrogen dioxide	sulphuric acid

   a) In car batteries.
   b) As a food preservative.
   c) To reduce the acidity of soil.
   d) For making detergents.

Challenge

1. A base accepts hydrogen ions (protons). In which equation is the underlined substance acting as a base?
   - **A** $2HCl + \underline{Ca} \longrightarrow CaCl_2 + H_2$
   - **B** $\underline{H_2O} + CO_2 \longrightarrow H_2CO_3$
   - **C** $\underline{N_2} + 3H_2 \longrightarrow 2NH_3$
   - **D** $NH_4^+ + \underline{OH^-} \longrightarrow NH_3 + H_2O$

2. Iron was added to dilute sulphuric acid. When the mixture was gently warmed, a gas was given off.
   a) Name and write the formula of the **two** ions present in dilute sulphuric acid.
   b) i) Suggest why the mixture had to be gently warmed.
      ii) What gas was given off?
      iii) What would you **see** that shows a gas was being given off?

Chemistry Today

When baking soda or baking powder is used in baking pastries, a crucial reaction between an acid and a carbonate takes place. Baking soda contains sodium hydrogencarbonate which, when mixed with an acid, reacts to form carbon dioxide (CO_2):

sodium hydrogencarbonate + hydrogen ion $\longrightarrow$ sodium ion + water + carbon dioxide

$$NaHCO_3(s) + H^+(aq) \longrightarrow Na^+(aq) + H_2O(l) + CO_2(g)$$

Baking soda and an acidic ingredient such as yoghurt, sour cream, lemon juice or vinegar must be mixed into pastry batter. Carbon dioxide is then released, gets trapped in the batter, and causes the batter to rise. The reaction produces light and delicious pastries.

When making éclairs and cream puffs, it is important that the gas is given off very quickly. For making this type of pastry, ammonium hydrogencarbonate or ammonium carbonate is used instead of sodium hydrogencarbonate. Ammonium substances decompose rapidly when heated, giving off carbon dioxide, ammonia and water.

ammonium hydrogencarbonate $\longrightarrow$ carbon dioxide + ammonia + water

$$NH_4HCO_3(s) \longrightarrow CO_2(g) + NH_3(g) + H_2O(l)$$

You can smell the ammonia given off during the cooking process. It is gone by the time the pastry is cooked. Ammonium hydrogencarbonate and ammonium carbonate are rarely used in homes because they are unstable and they decompose even at room temperature. In fact, they used to have another use as smelling salts — in earlier times, the smell of ammonia helped to revive someone who had fainted!

CRITICAL THINKING

What do you think would happen if (a) too little and (b) too much baking soda were used in a recipe?

Chapter 12

Salts

Your tears taste salty, your perspiration tastes salty, and your blood tastes salty. This is not just because you eat common salt or sodium chloride every day. Other compounds in your body are also known as salts and give that 'salty taste' to your blood, sweat and tears. Iron(II) sulphate is a salt that helps your blood carry oxygen. Sodium fluoride is a salt that stops your teeth from decaying. In this chapter, you will learn about different salts and how to make them in the school laboratory.

Chapter Outline

12.1 | Salts

What are salts?

All salts are ionic compounds. A **salt** is formed when *a metallic ion or an ammonium ion (NH_4^+) replaces one or more hydrogen ions of an acid.* For example, zinc sulphate is a salt. It is formed in the following reaction.

zinc hydroxide + sulphuric acid ⟶ zinc sulphate + water

$$Zn(OH)_2(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + 2H_2O(l)$$

Zn^{2+} comes from the base, $Zn(OH)_2$ — $Zn\ SO_4$ — SO_4^{2-} comes from sulphuric acid, H_2SO_4

A salt contains a positive metal ion (or ammonium ion) and a negative non-metal ion.

The following are some examples of salts made from acid-base reactions (Table 12.1).

Salt	Chemical formula	Possible reactants	
		Base	Acid
calcium chloride	$CaCl_2$	CaO	HCl
ammonium sulphate	$(NH_4)_2SO_4$	aqueous NH_3	H_2SO_4
copper(II) nitrate	$Cu(NO_3)_2$	$CuCO_3$	HNO_3

Table 12.1 Salts made from acid-base reactions

Quick Check

What acids are used to make the following salts?
a) Potassium nitrate
b) Copper(II) chloride
c) Sodium sulphate
d) Calcium phosphate
e) Ammonium ethanoate

Water of Crystallisation

Many salts combine with water molecules to form **crystals**. These water molecules are known as **water of crystallisation**.

Salts that contain water of crystallisation are called **hydrated salts**. *Salts that do not contain water of crystallisation* are called **anhydrous salts**. Anhydrous salts are often powders. Table 12.2 shows the formulae of some anhydrous salts and some hydrated salts. Compare the formulae of anhydrous and hydrated salts. How do they differ?

Name of salt	Formula of anhydrous salt	Formula of hydrated salt
copper(II) sulphate	$CuSO_4$	$CuSO_4.5H_2O$
magnesium sulphate	$MgSO_4$	$MgSO_4.7H_2O$
sodium carbonate	Na_2CO_3	$Na_2CO_3.10H_2O$
zinc sulphate	$ZnSO_4$	$ZnSO_4.7H_2O$

Table 12.2 Anhydrous and hydrated salts

The dot '.' in a formula means that if the substance is heated, everything after the dot will be given off first. For hydrated salts, water will be given off first.

Chem-Aid

Remember, crystals or hydrated salts are perfectly dry!

How can water of crystallisation be removed from a salt?
When a hydrated salt is heated, water of crystallisation is given off.

hydrated copper(II) sulphate (blue) $\xrightarrow{\text{heat}}$ anhydrous copper(II) sulphate (white) + water

$$CuSO_4.5H_2O \xrightarrow{\text{heat}} CuSO_4 + 5H_2O$$

Soluble and Insoluble Salts

Although salts are ionic compounds, not all salts are soluble in water. Figure 12.1 summarises the solubilities of the common salts in water at room temperature.

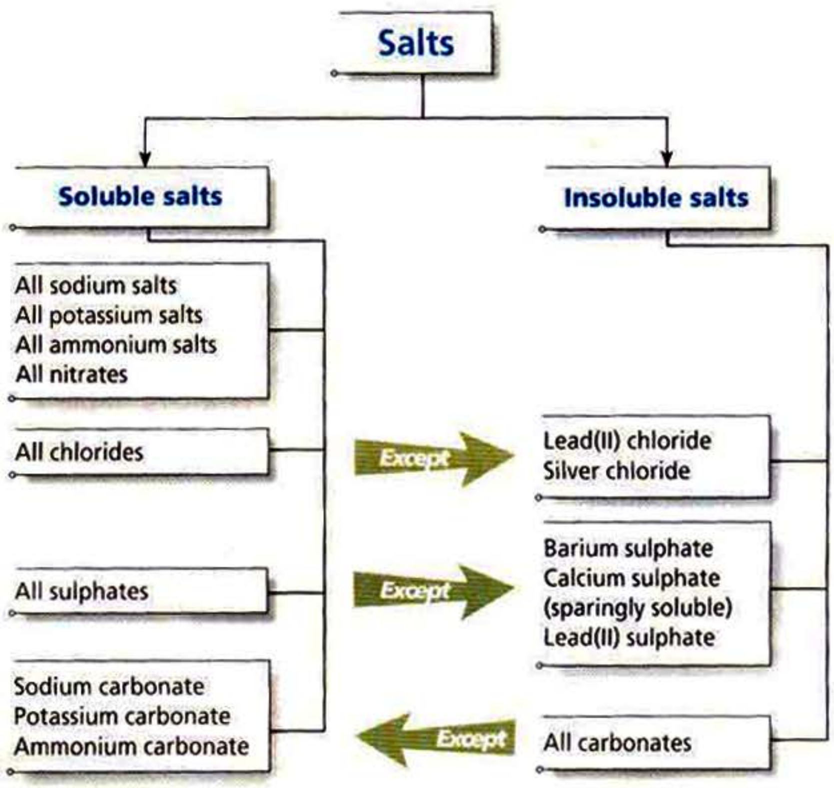


Fig. 12.1 The solubilities of salts in water

Quick Check

Are these salts soluble in water?
a) Ammonium phosphate
b) Barium carbonate
c) Chromium sulphate
d) Copper(II) nitrate
e) Nickel(II) chloride

12.2 | Preparing Salts

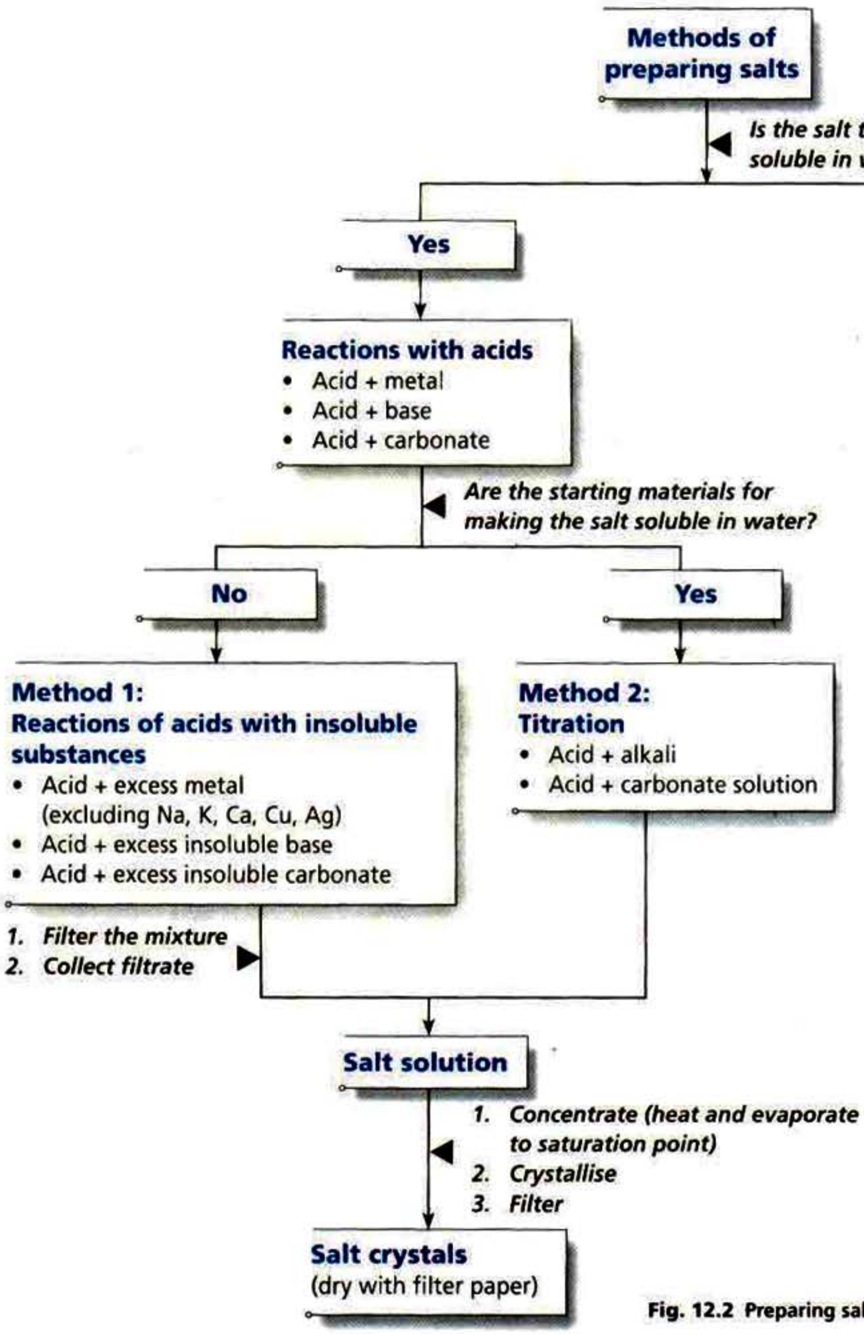
There are several ways to make salts. Before deciding on how to prepare a salt, you must consider two factors:
1. Is the salt soluble in water?
2. Are the starting materials soluble in water?

If the salt is soluble, it is usually prepared from reactions with acids. If the salt is insoluble, it is usually prepared by **precipitation** method. Fig. 12.2 summarises the methods of preparing salts in the laboratory.

Link

Reactions of acids were covered in chapter 11.

Learn more about metals in chapter 14.



Chem-Aid

A base is a metal oxide or a metal hydroxide.

Fig. 12.2 Preparing salts in the laboratory

TidBit

Zinc sulphate, sodium nitrate and copper(II) sulphate are needed to produce fertilisers. Magnesium chloride is used to make artificial leather. Lead(II) sulphate was previously used in paint but it has been banned as it is toxic.

You will now learn how to prepare
- zinc sulphate, copper(II) sulphate, magnesium chloride and sodium nitrate by reactions with acids,
- lead(II) sulphate by precipitation.

Method 1: Reacting an Acid with an Insoluble Metal, Base or Carbonate

As you have learnt in the previous chapter, acids react with the following substances:
1. Reactive metals such as magnesium, aluminum, zinc and iron
2. Bases (metal oxides or hydroxides)
3. Carbonates

In method 1, we react the acid with an **excess** of the substance. This ensures that all the acid is used up. The substance must also be **insoluble** in water. This allows the excess substance to be filtered from the salt solution produced.

How can we prepare a salt by reacting an acid with a metal?
For example, to prepare zinc sulphate, we react zinc with dilute sulphuric acid.

zinc + sulphuric acid $\longrightarrow$ zinc sulphate + hydrogen

$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$

Excess zinc must be used to prepare zinc sulphate. This makes sure that all the sulphuric acid is used up. Otherwise, the salt produced will be contaminated with sulphuric acid.

Chem-Aid

1. For method 1, it is crucial to use an excess of the solid.
2. Effervescence is the release of gas bubbles due to chemical change in a liquid solution. For example, when magnesium is added to hydrochloric acid, effervescence is observed as hydrogen is released.

The main steps involved in the preparation of zinc sulphate from zinc metal are shown below.

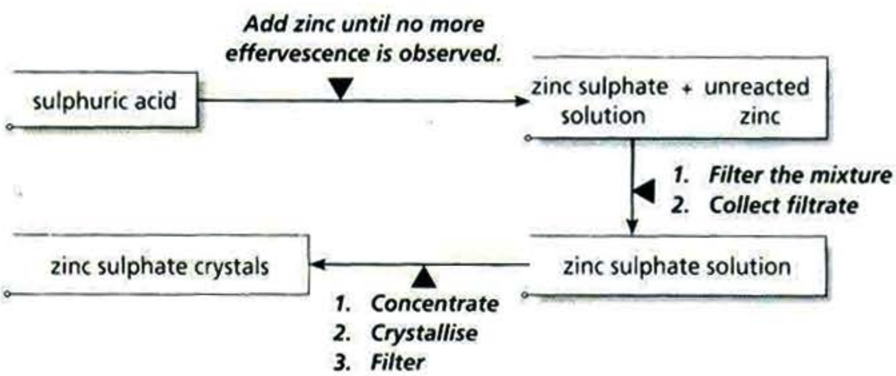


Fig. 12.3 Preparation of zinc sulphate

How can we prepare zinc sulphate from hydrochloric acid and zinc metal?

The preparation, separation and purification of zinc sulphate from zinc powder are shown below.

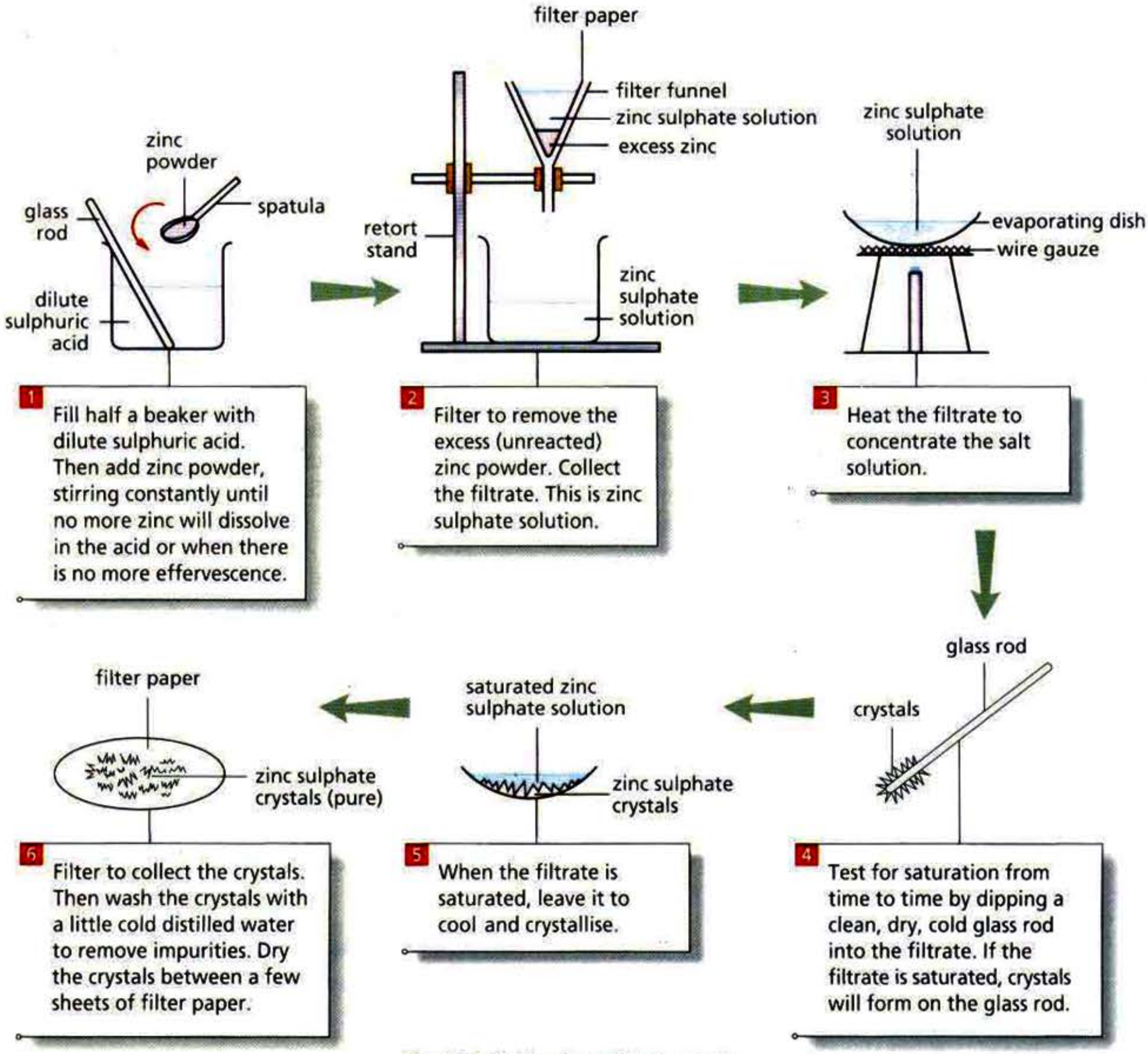


Fig. 12.4 Making zinc sulphate crystals

Why is this method NOT suitable for some metals?

Method 1 is suitable for moderately reactive metals such as zinc, magnesium, aluminium and iron. However, it is *not* suitable for

- very reactive metals such as potassium, sodium and calcium. These metals react violently with acid, so the reaction is very dangerous.
- unreactive metals such as copper and silver. These metals do not react with dilute acids.

Quick Check

Which of these salts can be made by reacting a metal with an acid?

a) Copper(II) chloride
b) Iron(II) chloride
c) Lead(II) chloride
d) Sodium chloride

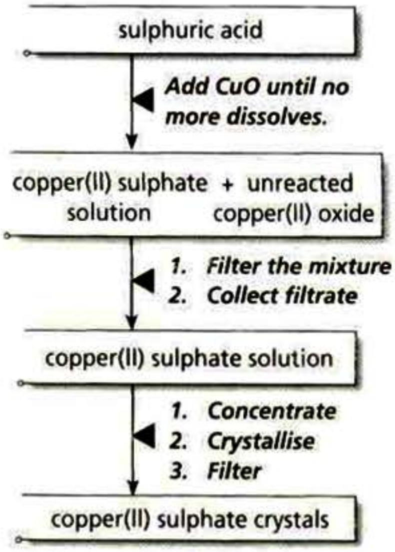


Fig. 12.5 Preparation of copper(II) sulphate

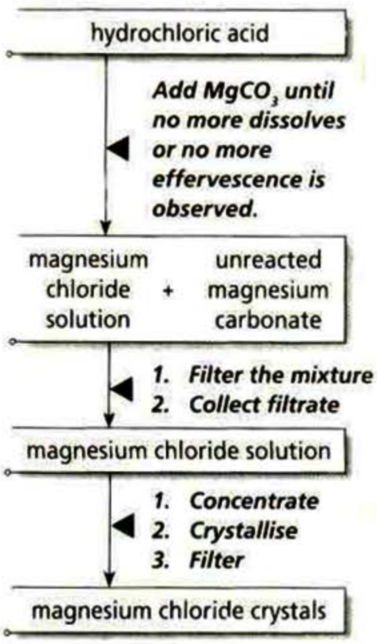


Fig. 12.6 Preparation of magnesium chloride

Chem-Aid

1. An alkali is a soluble base.
2. Examples of indicators are methyl orange, screened methyl orange and phenolphthalein.

How can we prepare a salt by reacting an acid with a base?

A salt can also be prepared by the reaction between an acid and a base (metal oxide or metal hydroxide).

When a metal is unreactive, it is better to react the metal oxide or hydroxide with an acid. For example, to produce copper(II) sulphate, we cannot use copper metal as it will not react with dilute sulphuric acid. However, copper(II) oxide will.

copper(II) oxide + sulphuric acid ⟶ copper(II) sulphate + water

$$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$$

Fig.12.5 shows the steps involved in the preparation of copper(II) sulphate crystals from an insoluble base, copper(II) oxide. The apparatus and procedure used are similar to that shown in Fig. 12.4.

How can we prepare a salt by reacting an acid with a carbonate?

We can also prepare soluble salts from insoluble carbonates. The steps are the same as those described for metal and bases. However, carbonates react quickly with acids, therefore no heating is required.

For example, magnesium chloride can be made by reacting dilute hydrochloric acid with magnesium carbonate.

magnesium carbonate + hydrochloric acid ⟶ magnesium chloride + carbon dioxide + water

$$MgCO_3(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$$

Excess solid (magnesium carbonate) is used to ensure all the hydrochloric acid is used up. When there is no more effervescence, all the acid has been used up. The excess solid can then be removed by filtration.

Fig. 12.6 shows the main stages in making magnesium chloride. The apparatus and procedure used are similar to that shown in Fig. 12.4.

Method 2: Titration

In method 1, we can tell when all the acid is used up. This is because we use an excess of the metal, base or carbonate that is insoluble in water. When all the acid is used up, we see a residue.

However, if we react the acid with a soluble substance (e.g. sodium hydroxide or sodium carbonate), no residue can be seen. How can we tell when the acid has been completely used up? To do so, we need to use an indicator to tell us exactly how much alkali or carbonate is required to neutralise the acid completely. This method is called **titration**. Sodium, potassium and ammonium salts are prepared using the titration method.

The main steps involved in the preparation of sodium nitrate are shown below.

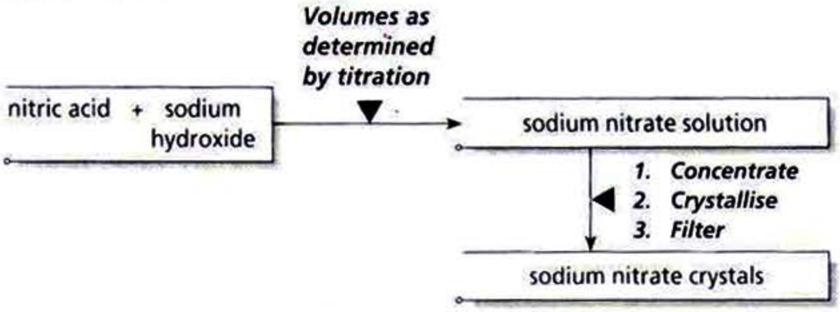


Fig. 12.7 Preparation of sodium nitrate

Experiment 1

The preparation of sodium nitrate involves
a) titration to determine the volumes of reactants required,
b) actual preparation of sodium nitrate.

a) Titration

1. Fill up a burette with dilute nitric acid. Note the initial burette reading (V_1 cm^3).
2. Pipette 25.0 cm^3 of dilute sodium hydroxide solution into a conical flask.
3. Add one or two drops of methyl orange (indicator) to the sodium hydroxide solution. The solution turns yellow.
4. Add dilute nitric acid from the burette slowly until the solution just turns orange permanently. This is the *end-point* (Fig. 12.8).
5. Record the final burette reading (V_2 cm^3). Hence, the volume of acid required for complete neutralisation = $(V_2 - V_1)$ cm^3.

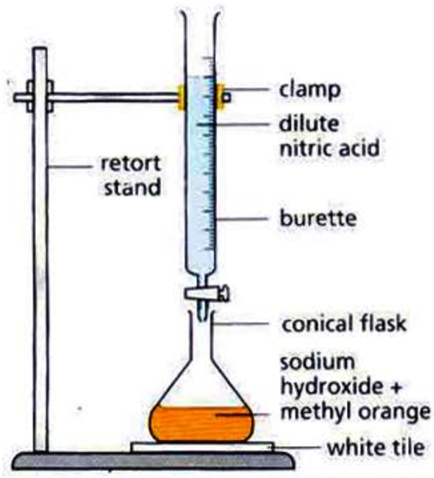


Fig. 12.8 Titration of sodium hydroxide solution with dilute nitric acid

b) Making sodium nitrate

1. Pipette 25.0 cm^3 of sodium hydroxide solution into a beaker. Then add $(V_2 - V_1)$ cm^3 of dilute nitric acid from the burette. (**Note**: Do not add the indicator, as it will make the salt impure.)
2. Heat the solution to evaporate the water until it is saturated.
3. Allow the saturated solution to cool so that the salt can crystallise.
4. Filter to collect the crystals.
5. Dry the crystals between a few sheets of filter paper.

Detecting End-point Using a pH Meter

Titration can also be carried out using a pH meter to detect the end-point of titration (Fig. 12.9). Dilute nitric acid is added slowly. A glass rod is used to stir the solution after each addition of dilute nitric acid.

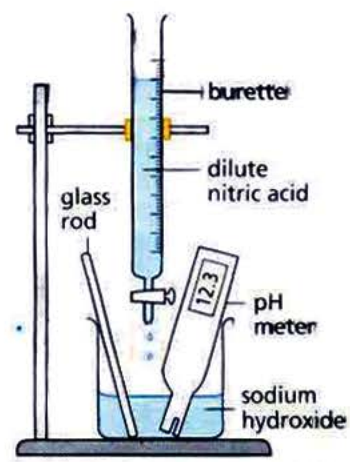


Fig. 12.9 Detecting the end-point of titration using a pH meter

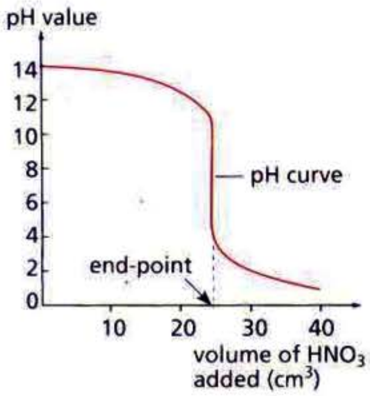


Fig. 12.10 Changes in pH value during an acid-base titration

The pH of the solution in the conical flask changes during an acid-base titration as solution is added from the burette. These changes in the pH values can be measured using a pH meter.

In the titration of a 1.0 mol/dm³ sodium hydroxide solution with a 1.0 mol/dm³ nitric acid, the pH decreases from 14 as more nitric acid is added. It finally ends at a pH value near 0. When the values of pH are plotted against the volume of nitric acid used, a pH curve is obtained (Fig. 12.10).

Chem-Aid

Because all nitrates are soluble in water, they are often used in preparing insoluble salts by precipitation.

Method 3: Precipitation

Table 12.3 shows some common salts that are insoluble in water.

Insoluble salts		
barium sulphate calcium sulphate lead(II) sulphate	lead(II) chloride lead(II) iodide silver chloride	all carbonates except carbonates of sodium, potassium and ammonium

Table 12.3 Examples of insoluble substances

All insoluble salts can be prepared by precipitation. To precipitate an insoluble salt, mix a solution that contains the positive ions of the salt with another solution that contains the negative ions of the salt to be prepared. For example, insoluble lead(II) sulphate can be prepared by using a soluble lead(II) salt (such as lead(II) nitrate) and dilute sulphuric acid or any soluble sulphate.

All lead salts are poisonous.

The preparation, separation and purification of lead(II) sulphate are shown below.

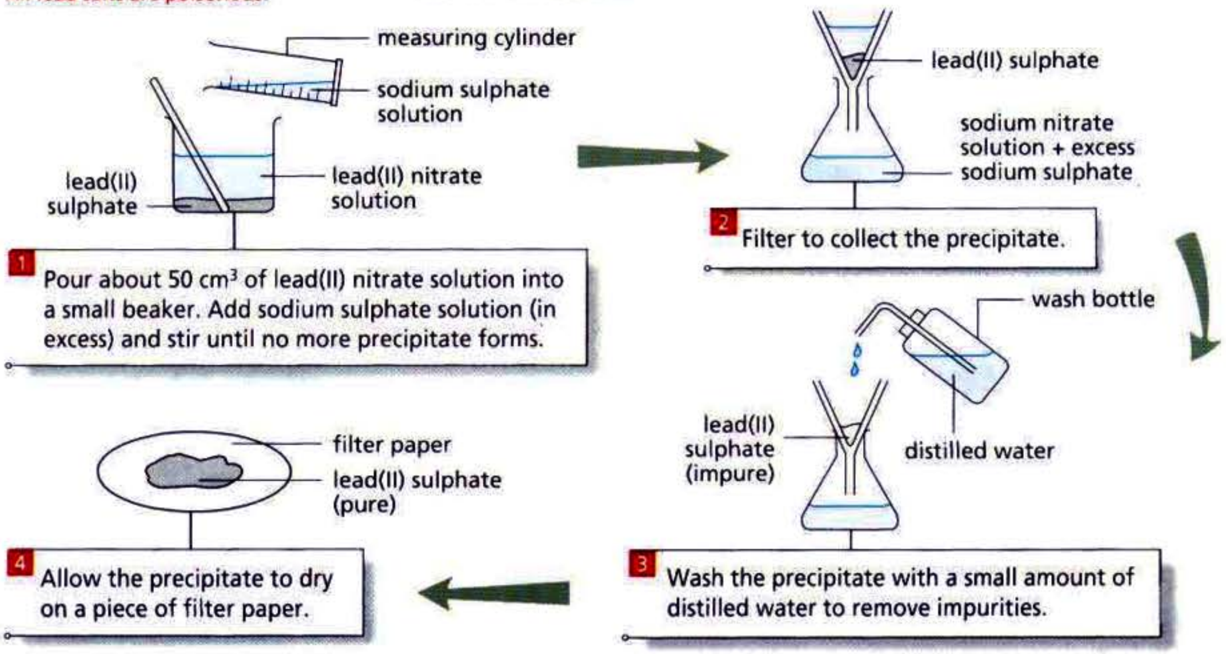


Fig. 12.11 Preparing lead(II) sulphate

Key ideas

1. Soluble salts are prepared by the following methods:
   - Acid + a metal (excluding potassium, sodium, calcium, copper and silver)
   - Acid + an insoluble base
   - Acid + an insoluble carbonate
   - Acid + an alkali (titration method)
2. Insoluble salts are prepared by the precipitation reaction of two soluble salt solutions.

Test Yourself 12.1

Worked Example 1

Which pair of reactants would not produce magnesium chloride?

A Magnesium and dilute hydrochloric acid.
B Magnesium carbonate and dilute hydrochloric acid.
C Magnesium nitrate and dilute hydrochloric acid.
D Magnesium oxide and dilute hydrochloric acid.

Thought Process

Magnesium nitrate and hydrochloric acid do not react, therefore they cannot be used to produce the soluble salt, magnesium chloride. Reactions A, B and D will produce magnesium chloride. A is the reaction between an acid and a reactive metal. B is the reaction of an acid with a carbonate. D is the reaction between an acid and a base.

Answer

C

Worked Example 2

Describe how you would prepare lead(II) iodide if you are given the following starting materials: lead(II) carbonate, dilute nitric acid and aqueous potassium iodide.

Thought Process

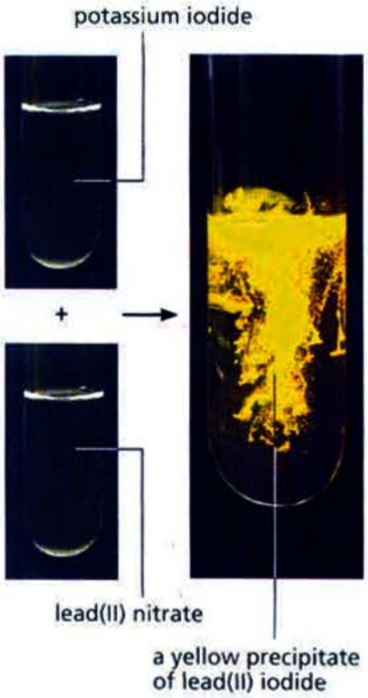
Lead(II) iodide is insoluble in water, so it must be prepared by the precipitation method. However, the starting material, lead(II) carbonate, is also insoluble in water. Hence, the first step is to convert lead(II) carbonate into a soluble lead(II) salt.

Answer

1. Add excess lead(II) carbonate to dilute nitric acid i.e. until no more effervescence is observed.

$$PbCO_3(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

2. Filter to remove excess lead(II) carbonate. The filtrate is lead(II) nitrate solution.



This reaction produced lead(II) iodide precipitate.

3. Add excess aqueous potassium iodide to the filtrate to precipitate lead(II) iodide.

$$2KI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$$

4. Filter the mixture to collect lead(II) iodide and wash the residue with distilled water. Allow the precipitate to dry on a piece of filter paper.

Questions

1. Identify substances *A* to *F* from the following reactions:

   calcium oxide + *A* ⟶ calcium nitrate + *B*
   C + *D* ⟶ zinc sulphate + hydrogen
   E + sodium carbonate ⟶ sodium chloride + *F* + water

2. Name the chemicals and state the key steps in preparing a pure sample of magnesium sulphate crystals.

3. Briefly describe how you would prepare a sample of potassium sulphate solution in the school laboratory.

4. | |
   |---|
   | lead(II) carbonate, ammonium sulphate, iron(II) sulphate, copper(II) nitrate, lithium chloride |

   For each of the above salts, name the chemicals and method you would use to prepare the salt.

12.3 | Qualitative Analysis

All salts are ionic compounds. Hence, they contain cations (positive ions) and anions (negative ions). We can make use of this idea to identify an unknown salt. In order to identify an unknown salt, it must be dissolved in water. We can then identify the cations and anions in the salt by adding certain reagents (chemicals) to separate portions of the salt solution. *This process of identification of cations and anions* is known as **qualitative analysis**.

Identifying Cations

Cations can be identified by using sodium hydroxide solution and aqueous ammonia. All cations (except Na^+, K^+ and NH_4^+) give precipitates with these alkalis.

A cation can be identified by noting
a) the colour of the precipitate produced,
b) whether the precipitate is soluble or insoluble in an excess of the reagent,
c) whether ammonia gas is liberated on addition of sodium hydroxide solution.

Chem-Aid

1. Aqueous ammonia is also known as ammonia solution. Its symbol is $NH_3(aq)$.
2. If there is no precipitate produced and no gas liberated, we may say that there is no visible reaction.

Table 12.4 summarises what we would observe when sodium hydroxide solution and ammonia solution are added separately to solutions containing different cations.

	Sodium hydroxide solution, NaOH(aq)	Aqueous ammonia, $NH_3(aq)$
	Observations on adding **a) a few drops of sodium hydroxide** **b) excess sodium hydroxide**	**Observations on adding** **a) a few drops of aqueous ammonia** **b) excess aqueous ammonia**
Al^{3+}	a) A white precipitate is formed. b) The precipitate dissolves in excess to give a colourless solution.	a) A white precipitate is formed. b) The precipitate is insoluble in excess.
Ca^{2+}	a) A white precipitate is formed. b) The precipitate is insoluble in excess.	a) No precipitate. b) No precipitate.
Cu^{2+}	a) A light blue precipitate is formed. b) The precipitate is insoluble in excess.	a) A light blue precipitate is formed. b) The precipitate dissolves in excess to give a deep blue solution.
Fe^{2+}	a) A green precipitate is formed. b) The precipitate is insoluble in excess.	a) A green precipitate is formed. b) The precipitate is insoluble in excess.
Fe^{3+}	a) A reddish-brown precipitate is formed. b) The precipitate is insoluble in excess.	a) A reddish-brown precipitate is formed. b) The precipitate is insoluble in excess.
Pb^{2+}	a) A white precipitate is formed. b) The precipitate dissolves in excess to form a colourless solution.	a) A white precipitate is formed. b) The precipitate is insoluble in excess.
Zn^{2+}	a) A white precipitate is formed. b) The precipitate dissolves in excess to form a colourless solution.	a) A white precipitate is formed. b) The precipitate dissolves in excess to form a colourless solution.
NH_4^+	a) No precipitate is formed. b) On heating, ammonia gas is given off. Ammonia turns moist red litmus paper blue.	

Table 12.4 Identifying cations

Why do we get the above observations?

The precipitate in each of the reactions is the hydroxide of the metal ion. For example, the reddish-brown precipitate seen on adding sodium hydroxide or aqueous ammonia to a solution containing iron(III) ions is iron(III) hydroxide, $Fe(OH)_3$.

iron(III) ion + hydroxide ion ⟶ iron(III) hydroxide

$$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$$

You would have noticed some of the precipitates dissolve in excess sodium hydroxide or aqueous ammonia (for example, solutions containing zinc ions or copper(II) ions). This is due to the formation of compounds that are soluble in water.

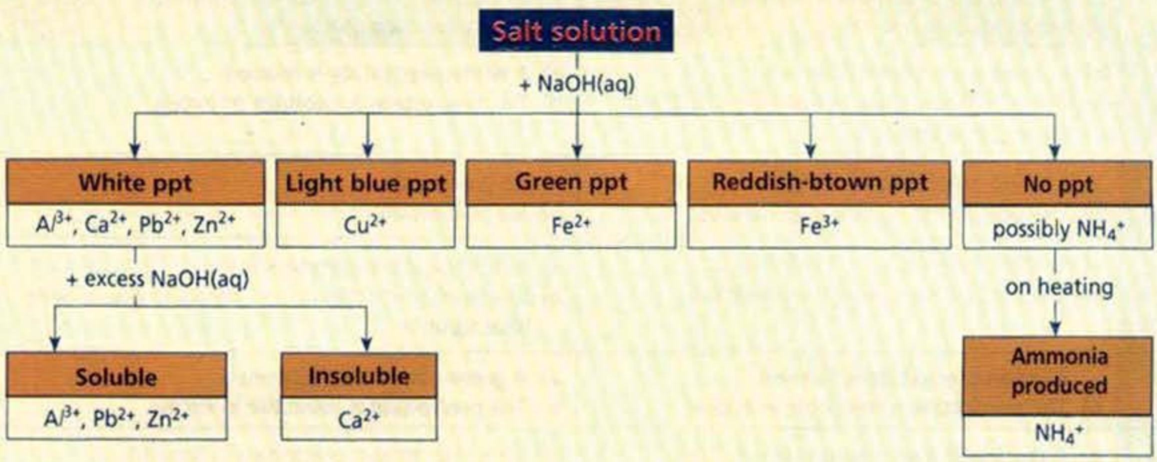
TidBit

Have you ever seen fireworks? The colour of each explosion is actually unique to a cation. Thus, we can also identify cations through the colour of their flames! Chemists call such identification methods 'flame tests'.

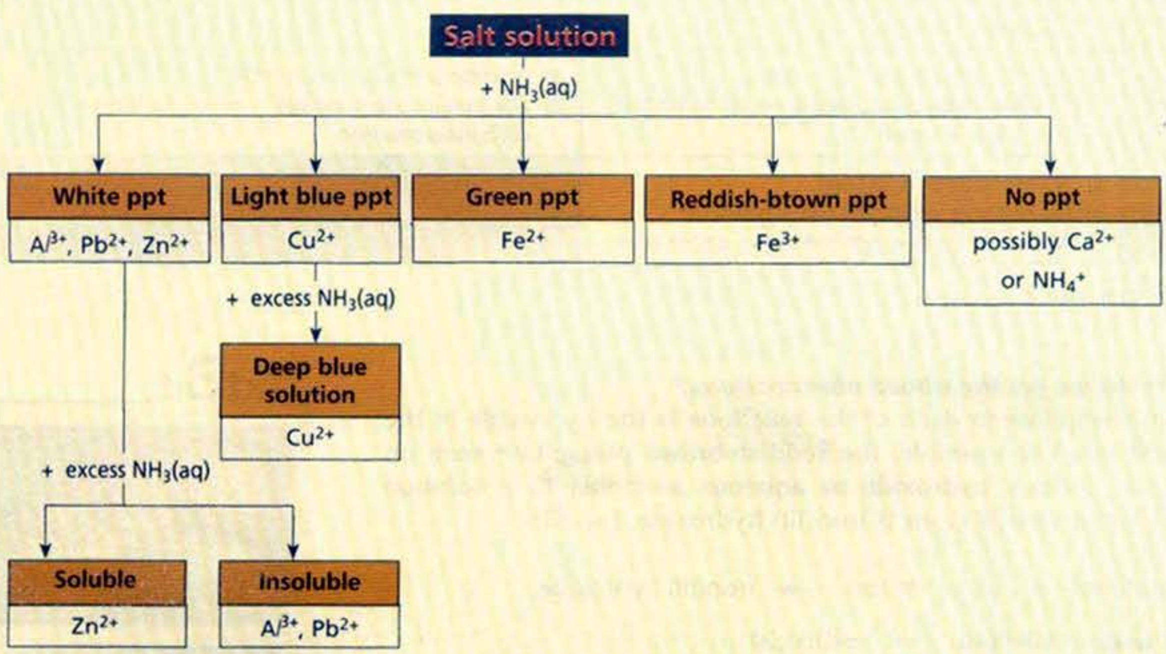
Key ideas

Use the following to identify the cations present in a salt:
(Note: ppt = precipitate)

1. Sodium hydroxide solution



2. Aqueous ammonia



Identifying Anions

The reactions shown in Table 12.5 can be used to identify some common anions.

Anion	Test	Observations and inference
CO_3^{2-} (carbonate)	Add dilute hydrochloric acid. Pass the gas given off into limewater.	Effervescence is observed. Gas given off forms a white precipitate with limewater. Carbon dioxide gas is given off.
Cl^- (chloride)	Add dilute nitric acid, then add **silver nitrate** solution.	A white precipitate of silver chloride is formed.
I^- (iodide)	Add dilute nitric acid, then add **silver nitrate** solution.	A yellow precipitate of silver iodide is formed.
NO_3^- (nitrate)	Add dilute sodium hydroxide. Then add a piece of aluminium foil. Warm the mixture. Test the gas given off with a piece of moist red litmus paper.	The moist red litmus paper turns blue. Ammonia gas is given off.
SO_4^{2-} (sulphate)	Add dilute nitric acid, then add **barium nitrate** solution.	A white precipitate of barium sulphate is formed.

Table 12.5 Identifying anions

Test Yourself 12.2

Worked Example

Barium nitrate solution is added to a solution of a salt. A white precipitate is formed. Which ion is present in the salt?

A Cl^-
B Ca^{2+}
C NH_4^+
D SO_4^{2-}

Thought Process

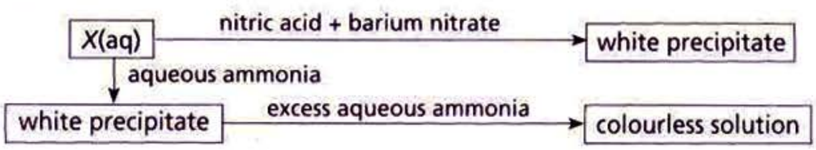
The white precipitate is barium sulphate ($BaSO_4$). The other ions do not form a precipitate with barium nitrate.

Answer

D

Question

Deduce the identity of the cation and anion present in salt *X* from the following observations:



Identifying Gases and Water

How do we identify gases?

A gas is often liberated (given off) when an unknown salt is being tested. Table 12.6 shows the tests used to identify common gases.

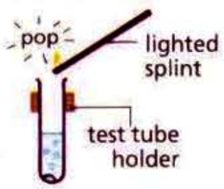
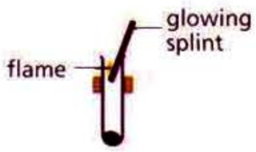
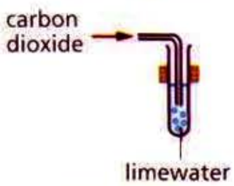
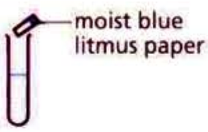
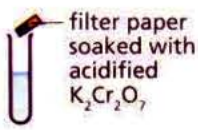
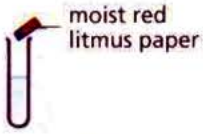
	Gas	Colour and odour	Diagram	Test	Observations
Odourless gases	Hydrogen	Colourless and odourless		Place a lighted splint at the mouth of the test tube.	The lighted splint is extinguished with a 'pop' sound.
	Oxygen	Colourless and odourless		Insert a glowing splint into the test tube.	The glowing splint is rekindled (i.e. catches fire).
	Carbon dioxide	Colourless and odourless		Bubble gas through limewater.	A white precipitate is formed.
Pungent gases	Chlorine	Greenish-yellow gas with a pungent smell		Place a piece of moist blue litmus paper at the mouth of the test tube.	The moist blue litmus paper turns red, and is then bleached.
	Sulphur dioxide	Colourless gas with a pungent smell		Place a piece of filter paper soaked with acidified potassium dichromate(VI) at the mouth of the test tube.	The orange potassium dichromate(VI) turns green.
	Ammonia	Colourless gas with a pungent smell		Place a piece of moist red litmus paper at the mouth of the test tube.	The moist red litmus paper turns blue.

Table 12.6 Identifying gases

How do we test for the presence of water?

When a substance is heated, a colourless liquid that condenses near the top of the test tube is most likely to be water given off by a hydrated salt. Pure water is colourless, odourless and tasteless. There are two chemical tests that will show whether a liquid is water.

1. **Test with anhydrous copper(II) sulphate**
   Water will change the colour of anhydrous copper(II) sulphate from white to blue.

   anhydrous copper(II) sulphate (white) + water → hydrated copper(II) sulphate (blue)

   $$CuSO_4(s) + 5H_2O(l) \longrightarrow CuSO_4.5H_2O(s)$$

2. **Test with anhydrous cobalt(II) chloride**
   Water will change the colour of dry cobalt(II) chloride paper from blue to pink.

   anhydrous cobalt(II) chloride (blue) + water → hydrated cobalt(II) chloride (pink)

   $$CoCl_2(s) + 6H_2O(l) \longrightarrow CoCl_2.6H_2O(s)$$

Note that these two tests only show the presence of water. They cannot be used to test for the purity of water.

Quick Check

How do you show that a sample of water is pure?

Key ideas

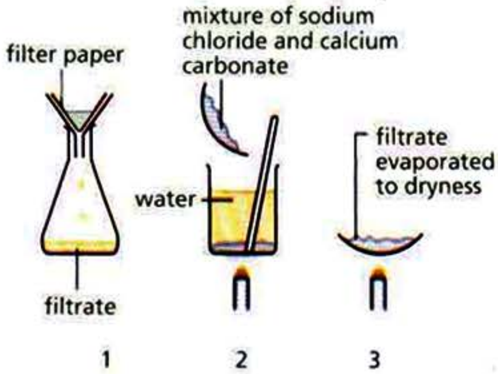
The table below shows the tests for different gases in a laboratory.

Observations	Identity of gas
Colourless, odourless gas that • extinguishes a lighted splint with a 'pop' sound.	• Hydrogen
• rekindles a glowing splint.	• Oxygen
• forms a white precipitate with limewater.	• Carbon dioxide
Colourless gas with a pungent smell that • turns acidified potassium dichromate(VI) from orange to green.	• Sulphur dioxide
• turns moist red litmus paper blue.	• Ammonia
Greenish-yellow gas with a pungent smell that turns moist blue litmus paper red, then bleaches it.	Chlorine

Exercise 12

Foundation

1. A mixture of sodium chloride and calcium carbonate is separated using the steps below. In which order should the steps be carried out?



   A 1, 2, then 3.
   B 2, 1, then 3.
   C 1, 3, then 2.
   D 3, 1, then 2.

2. When aqueous sodium hydroxide and aluminium foil are warmed with substance *X*, an alkaline gas is given off. What anion does *X* contain?

   A Carbonate **B** Chloride
   C Nitrate **D** Sulphate

3. Which substance will **not** form magnesium sulphate when reacted with dilute sulphuric acid?

   A Magnesium carbonate
   B Magnesium hydroxide
   C Magnesium nitrate
   D Magnesium powder

4. What happens when potassium chloride solution is added to silver nitrate solution?

   A One soluble salt and one insoluble salt are formed.
   B There is no reaction.
   C Two insoluble salts are formed.
   D Two soluble salts are formed.

5. A farmer heated some fertiliser with sodium hydroxide solution. A gas that was given off turned moist red litmus paper blue. He then dissolved some of the fertiliser in water. On adding hydrochloric acid and barium chloride, a white precipitate was formed. What salt did his fertiliser contain?

   A Ammonium nitrate
   B Ammonium sulphate
   C Potassium nitrate
   D Potassium sulphate

6. A solution *S* contains dilute nitric acid and lead(II) nitrate. Which set of results would be obtained if separate samples of *S* were tested with Universal Indicator solution and aqueous potassium iodide?

	Universal Indicator pH	Potassium iodide(aq)
A	3	yellow precipitate
B	5	no reaction
C	7	no reaction
D	9	yellow precipitate

7. a) The following table shows how salts can be prepared. Identify the missing chemicals *X*, *Y* and *Z* in the table.

Suitable chemicals	Salt formed
i) iron + *X*	iron(II) sulphate
ii) sulphuric acid + *Y*	sodium sulphate
iii) sulphuric acid + *Z*	copper(II) sulphate

   b) i) Name the reagent used with aluminium to identify the nitrate ion.
   ii) What condition is needed for the reaction to take place?
   iii) How is the gas given off identified?

   c) Describe how you would prepare a dry sample of lead(II) sulphate starting from lead(II) oxide.

8. a) Describe **four** ways to test whether solution is acidic. For each test, explain the observation that proves that the solution is acidic.

   b) i) What would you observe when barium chloride solution is added to dilute sulphuric acid?
   ii) Write down the chemical equation for the reaction in (b)(i).

Challenge

1. There are two unlabelled test tubes. One contains sodium carbonate solution and the other contains sodium chloride solution. Which test could identify the solutions?
   - **A** Add aqueous ammonia.
   - **B** Add dilute hydrochloric acid.
   - **C** Add lead(II) nitrate solution.
   - **D** Add sodium hydroxide solution.

2. Which of the following is the best pair of reagents to use for preparing copper(II) chloride crystals?
   - **A** Copper metal and dilute hydrochloric acid.
   - **B** Copper(II) carbonate and aqueous ammonium chloride.
   - **C** Copper(II) sulphate and dilute hydrochloric acid.
   - **D** Copper(II) oxide and dilute hydrochloric acid.

3. When 25 cm^3 of dilute sulphuric acid was titrated against 0.1 mol/dm^3 sodium hydroxide solution using phenolphthalein as the indicator, the following results were obtained.

	Burette readings		
Titration number	1	2	3
Final reading (cm^3)	33.3	24.8	24.9
Initial reading (cm^3)	8.1	0.0	0.0
Volume of sodium hydroxide used (cm^3)			

   a) Complete the above table.

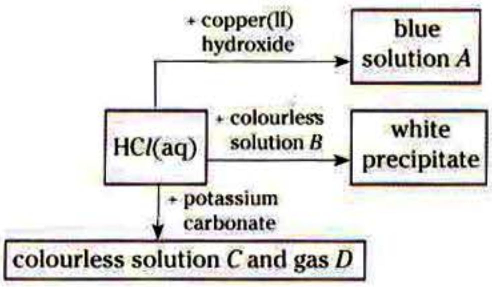
   b) Determine the average volume of sodium hydroxide required for this titration.

   c) Write the chemical equation for the reaction between sodium hydroxide and sulphuric acid.

   d) Calculate the number of moles of sulphuric acid used in the titration.

   e) Hence, calculate the concentration, in mol/dm^3, of the sulphuric acid.

4. A reaction scheme involving dilute hydrochloric acid is shown below. Give the names and formulae of *A*, *B*, *C* and *D*.



5. Describe **one** chemical test in each case to distinguish between the two chemicals.

   a) $NH_4Cl(aq)$ and $MgCl_2(aq)$

   b) $Zn(OH)_2(s)$ and $Mg(OH)_2(s)$

6. State in words the reactions represented by the following ionic equations.

   a) $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

   b) $Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

   c) $CuSO_4(s) \xrightarrow{+H_2O} Cu^{2+}(aq) + SO_4^{2-}(aq)$

7. What deductions can be made from the following observations on a crystalline compound *K*? Explain your answers.

   a) When heated gently, it gave off a colourless vapour that condensed on the side of the test tube. The liquid turned blue cobalt(II) chloride paper pink.

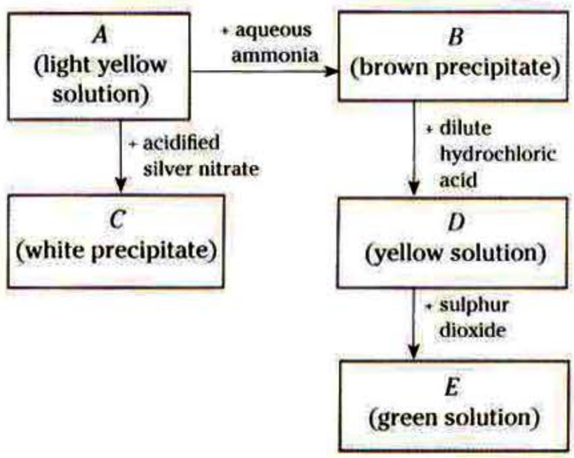
   b) *K* was dissolved in water. The solution formed was divided into three parts.

      i) *Part 1.* Sodium hydroxide solution was added and the mixture heated. A gas was given off that turned moist red litmus paper blue.

      ii) *Part 2.* Aqueous ammonia was added and a blue precipitate was formed. The precipitate was soluble in excess aqueous ammonia.

      iii) *Part 3.* The solution was acidified with dilute hydrochloric acid and then barium chloride solution was added. A white precipitate was formed.

8. Consider the flow chart as shown below.



   a) Identify substances *A* – *E*.

   b) Instead of using silver nitrate solution, suggest another substance that could be used for reaction *A* $\longrightarrow$ *C*.

Chapter 13
Oxidation and Reduction

Chapter Outline

Have you have ever walked along the Singapore River? Then you might have noticed these life-size sculptures of cats. If you take a very close look at the sculptures, you might notice that they have a green coating. The coating is the result of a chemical reaction, involving the processes of **oxidation** and **reduction**. In this chapter, you will learn more about oxidation, reduction and redox reactions.

13.1 | Oxidation Reactions

What is oxidation?

The term **oxidation** was first used to describe reactions in which oxygen combined with an element or a compound to form another substance. Nowadays, we say that *a substance is oxidised if it gains oxygen, loses hydrogen, loses electrons or increases its oxidation state after a reaction.*

Oxidation — Gain of Oxygen

When calcium burns in oxygen, the following reaction takes place:

calcium + oxygen ⟶ calcium oxide

$$2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$$

Since calcium has gained oxygen, we say that it has been oxidised to calcium oxide. This process is called oxidation.

Methane is an important fuel used in Bunsen burners in your school laboratory.

Oxidation also takes place when methane burns in oxygen. The equation for the reaction is

methane + oxygen ⟶ carbon dioxide + water vapour

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

The carbon atom in methane has gained oxygen and has been oxidised to carbon dioxide. The hydrogen atom in methane has also gained oxygen. It has been oxidised to water.

Oxidation — Loss of Hydrogen

A substance is oxidised if it loses hydrogen. When ammonia is passed over heated copper(II) oxide, the following reaction occurs:

ammonia + copper(II) oxide ⟶ nitrogen + copper + water vapour

$$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(g)$$

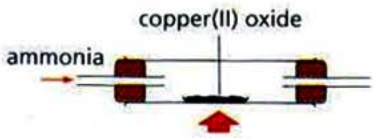


Fig. 13.1 Reaction of ammonia and copper(II) oxide

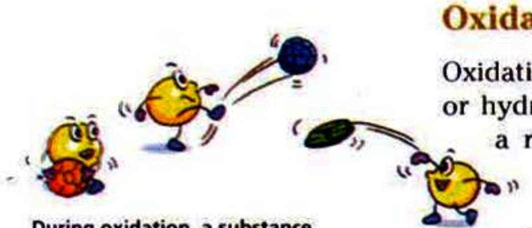
Ammonia has lost hydrogen. It has been oxidised to nitrogen.

Oxidation also takes place when hydrogen sulphide gas and chlorine gas are mixed. A yellow deposit of sulphur forms and a choking gas, hydrogen chloride, is produced.

hydrogen sulphide + chlorine ⟶ hydrogen chloride + sulphur

$$H_2S(g) + Cl_2(g) \longrightarrow 2HCl(g) + S(s)$$

Hydrogen sulphide has been oxidised to sulphur.



During oxidation, a substance receives oxygen, loses hydrogen or loses electrons.

Oxidation — Loss of Electrons

Oxidation reactions can take place even if there is no oxygen or hydrogen present. When a substance loses electrons during a reaction, it is oxidised. Therefore, in terms of electron transfer, oxidation is defined as *the loss of electrons from a substance*. Consider the reaction between magnesium and chlorine to form magnesium chloride.

magnesium + chlorine ⟶ magnesium chloride

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

Electrons have been transferred during this reaction. The following half equations represent the transfer of electrons.

$Mg(s) \longrightarrow Mg^{2+}(s) + 2e^-$ — **Magnesium atoms lose electrons to form magnesium ions.**

$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(s)$ — **Chlorine molecules gain electrons to form chloride ions.**

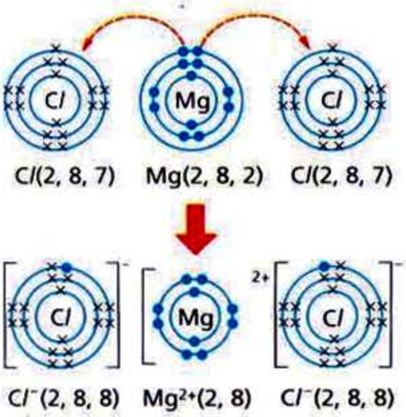


Fig. 13.2 The transfer of electrons during an oxidation reaction

In this reaction, magnesium has been oxidised.

What happens when copper is placed into silver nitrate solution?

If a coil of copper is placed in a solution of silver nitrate (Fig. 13.3), the copper slowly reacts and the solution turns blue. At the same time, the copper coil becomes coated with a layer of silver metal. The overall equation for the reaction is

copper + silver nitrate ⟶ copper(II) nitrate + silver

$Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

The ionic equation for this reaction is

$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$

In this reaction, each copper atom loses two electrons and becomes a copper(II) ion in aqueous solution.

Oxidation of copper has taken place. The copper atom has been oxidised to the copper(II) ion.

Link

Atoms of one metal can replace the ions of another metal in a salt solution. This is known as a metal displacement reaction. Find out more in chapter 14.

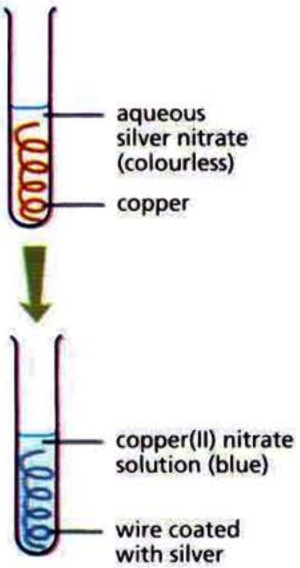


Fig. 13.3 The reaction between copper metal and silver nitrate solution

Oxidation — Increase in Oxidation State

The **oxidation state** is *the charge an atom of an element would have if it existed as an ion in a compound* (even if it is actually covalently bonded).

To work out the oxidation state of an atom, we apply the following rules:

Rule	Example	Oxidation state	
1. The oxidation state of a free element is zero.	Cu S Cl_2	0 0 0	
2. The oxidation state of a simple ion is the same as the charge on the ion.	K^+ Zn^{2+} Cl^- O^{2-}	+1 +2 −1 −2	
3. The oxidation states of the atoms present in the formula of a compound add up to zero. The example shown here is $CaCO_3$.	Ca C 3O	+2 +4 3 × (−2) = −6	Add up to zero
4. The total of the oxidation states of the atoms in a polyatomic ion is equal to the charge on the ion. The example shown here is SO_4^{2-}.	S 4O	+6 4 × (−2) = −8	Add up to −2

Table 13.1 Rules for working out the oxidation states of atoms and ions

Link

What are the charges on common ions? Recall from chapter 6.

Chem-Aid

1. Oxidation state is also known as oxidation number.
2. Some metals (known as transition elements)have variable oxidation states. Iron can exist as Fe^{2+} and Fe^{3+} in compounds. When Fe^{2+} reacts and forms Fe^{3+}, the oxidation state of iron has increased from +2 to +3.
3. Polyatomic = more than one atom

How can we determine the oxidation states of atoms in a compound?

In the compound magnesium oxide, MgO, magnesium exists as the ion Mg^{2+}. Since magnesium has two positive charges, its oxidation state is +2 (rule 2). Oxygen exists as the ion, O^{2-}, thus it has the oxidation state −2 (Fig. 13.4).

MgO

(+2 −2) Sum = 0

Fig. 13.4

Carbon dioxide, CO_2, is a covalent compound. However, we can assume each oxygen 'ion' has an oxidation state of −2 (rule 2). Using rule 3, carbon has an oxidation state of +4 (Fig. 13.5).

CO_2

(+4 −2 x 2) Sum = 0

Fig. 13.5

How can we use oxidation states to identify oxidation?

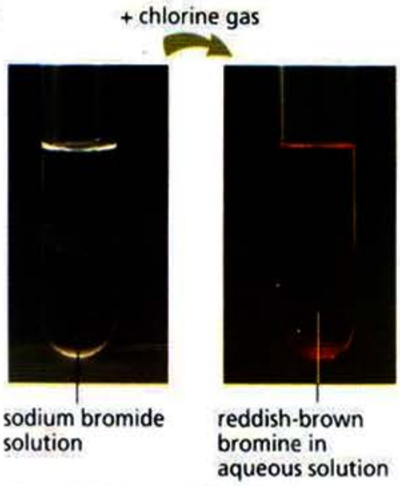
If the oxidation state of a substance increases, the substance has been oxidised. Consider the reaction between zinc and a solution of copper(II) ions.

increase in oxidation state

0 → +2

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

Zinc atoms (oxidation state = 0) become zinc ions, Zn^{2+} (oxidation state = +2). The oxidation state of zinc has increased from 0 to +2. We say that zinc has been oxidised to zinc ions.



When chlorine gas is bubbled into sodium bromide solution, reddish-brown bromine is produced. How can we tell which substance has been oxidised?

To check whether oxidation has taken place in any reaction, follow the three steps below:

1. Write the balanced equation for the reaction.
2. Write the oxidation states of all the substances in the reaction.
3. Compare the oxidation states to check which reactant has been oxidised.

For example, when chlorine gas is bubbled into sodium bromide solution, bromide ions, Br^- are displaced to form bromine molecules, Br_2. We know this because bromine molecules formed turned the solution reddish-brown. (Bromide ions are colourless.)

1. Write the balanced equation for the reaction.

$$Cl_2(g) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$$

2. Write the oxidation states of all the substances in the reaction.

$$\underset{\substack{\downarrow \\ 0}}{Cl_2(g)} + \underset{\substack{\downarrow \\ -1}}{2Br^-(aq)} \longrightarrow \underset{\substack{\downarrow \\ -1}}{2Cl^-(aq)} + \underset{\substack{\downarrow \\ 0}}{Br_2(aq)} \quad \text{oxidation state}$$

3. Compare the oxidation states to check which reactant has been oxidised . The bromide ions have been oxidised to aqueous bromine because there is an increase in the oxidation state from –1 to 0.

Key ideas

A substance is oxidised if it gains oxygen, loses hydrogen, loses electrons or increases its oxidation state after a reaction.

Test Yourself 13.1

Worked Example

What happens to magnesium when it is oxidised to $MgCl_2$?

It gains oxygen.
It loses two protons.
Its electronic configuration becomes (2, 8, 8).
Its oxidation state increases.

Thought Process

The reaction that takes place is $Mg \longrightarrow Mg^{2+}$. The magnesium loses twc electrons to become (2, 8). Its oxidation state increases from 0 to +2.

Answer

D

Questions

1. Identify the s ıbstances that are oxidised in the reactions below. Explain your answers.
   a) hydrogen + copper(II) oxide ⟶ copper + water
   b) lead(II) oxide + carbon monoxide ⟶ lead + carbon dioxide
   c) magnesium + hydrochloric acid ⟶ magnesium chloride + hydrogen

2. On heating concentrated hydrochloric acid with manganese(IV) oxide, the following reaction takes place.

$$4HCl + MnO_2 \longrightarrow 2H_2O + MnCl_2 + Cl_2$$

Which substance has been oxidised? Explain your answer.

13.2 | Reduction Reactions

What is reduction?

A reduction reaction is the reverse process of an oxidation reaction. **Reduction** has taken place if a substance *loses oxygen, gains hydrogen, gains electrons or decreases its oxidation state after a reaction.*

Reduction — Loss of Oxygen

When a mixture of zinc powder and copper(II) oxide is heated, the following reaction occurs:

zinc + copper(II) oxide ⟶ zinc oxide + copper

$$Zn(s) + CuO(s) \longrightarrow ZnO(s) + Cu(s)$$

In this reaction, the copper(II) oxide has lost its oxygen. It is reduced to copper metal.

$$Zn(s) + \underbrace{CuO(s) \longrightarrow ZnO(s) + Cu(s)}_{\text{reduced}}$$

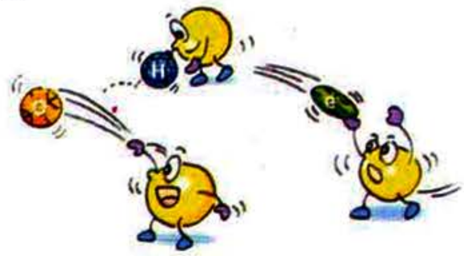
Reduction — Gain of Hydrogen

When a mixture of chlorine and hydrogen is exposed to sunlight, it explodes and produces white fumes of hydrogen chloride.

hydrogen + chlorine ⟶ hydrogen chloride

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

Chlorine is reduced as it has gained hydrogen.



During reduction, a substance loses oxygen, gains hydrogen or gains electrons.

Reduction — Gain of Electrons

Reduction is also defined as the *gain of electrons by a substance.* When hydrogen sulphide gas is passed into iron(III) chloride solution, a green solution of iron(II) chloride and a light yellow precipitate of sulphur are produced.

iron(III) chloride + hydrogen sulphide ⟶ iron(II) chloride + hydrogen chloride + sulphur

$$2FeCl_3(aq) + H_2S(g) \longrightarrow 2FeCl_2(aq) + 2HCl(aq) + S(s)$$

We can rewrite the above chemical equation in the form of an ionic equation.

$$2Fe^{3+}(aq) + H_2S(g) \longrightarrow 2Fe^{2+}(aq) + 2H^+(aq) + S(s)$$

(Note that the Cl^- ions are spectator ions.)

In this reaction, each iron(III) ion has gained an electron to form an iron(II) ion, i.e.

$$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$$

This means that iron(III) ions are reduced to iron(II) ions.

Reduction — Decrease in Oxidation State

Reduction can also be defined as *the decrease in the oxidation state of a substance after a reaction.*

The ionic equation for the reaction between copper(II) sulphate and zinc is

$$\underset{0}{\underset{\downarrow}{Zn(s)}} + \underset{+2}{\underset{\downarrow}{Cu^{2+}(aq)}} \longrightarrow \underset{+2}{\underset{\downarrow}{Zn^{2+}(aq)}} + \underset{0}{\underset{\downarrow}{Cu(s)}} \quad \text{oxidation state}$$

The oxidation state of copper has decreased by 2. The copper(II) ions, Cu^{2+}, have been reduced to copper atoms, Cu.

Key ideas

A substance is reduced if it loses oxygen, gains hydrogen, gains electrons or decreases its oxidation state after a reaction.

Test Yourself 13.2

Worked Example

Which reaction is an example of reduction?

A Copper(II) oxide to copper.
B Copper(II) oxide to copper(II) sulphate.
C Hydrochloric acid to chlorine.
D Iron(II) chloride to iron(III) chloride.

Thought Process

A Copper(II) oxide has lost oxygen to become copper, thus copper(II) oxide is reduced.

B The reaction is neither oxidation nor reduction as the oxidation state of the copper(II) ion was not changed.
C Hydrochloric acid has lost hydrogen to become chlorine, thus it is oxidised.
D Iron is oxidised because its oxidation state increased from +2 to +3.

Answer
A

Questions

1. Which substance is reduced in each of the following reactions?
   a) $PbO(s) + H_2(g) \longrightarrow Pb(s) + H_2O(l)$
   b) $H_2S(g) + Cl_2(g) \longrightarrow 2HCl(g) + S(s)$

2. Complete the table comparing oxidation and reduction.

Process	Oxygen	Hydrogen	Electron	Oxidation number
Oxidation	gain			increase
Reduction		gain		

13.3 | Redox Reactions

Oxidation and reduction always take place together. In other words, there can be no oxidation without reduction and vice versa. We call the combined process a **redox reaction**.

When steam is passed over heated magnesium, magnesium oxide and hydrogen gas are produced.

magnesium + steam ⟶ magnesium oxide + hydrogen

$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$

Has a redox reaction taken place?
A redox reaction has taken place because magnesium has gained oxygen and is oxidised to magnesium oxide. Water has lost oxygen and is reduced to hydrogen.

What happens when a black-and-white film is exposed to sunlight?
Black-and-white photographic film contains tiny crystals of silver bromide suspended in gelatine. When a photographic film is exposed to sunlight, the following redox reaction takes place:

$2Ag^+Br^-(s) \longrightarrow 2Ag(s) + Br_2(g)$

Why do old black-and-white photographs turn brown?

Link

Iron will oxidise when left out in the open. Find out in chapter 14 how iron is oxidised and what can be done to prevent it.

Chem-Aid

OIL RIG
Oxidation Is Loss Reduction Is Gain of electrons

Can you deduce which substance is reduced and which is oxidised?

Since silver ions have gained electrons to form silver metal, the silver ions are reduced. Bromide ions have lost electrons and are oxidised to bromine molecules.

$Ag^+(s) + e^- \longrightarrow Ag(s)$ (reduction)

$2Br^-(s) \longrightarrow Br_2(g) + 2e^-$ (oxidation)

$$\overbrace{2Ag^+Br^-(s) \longrightarrow 2Ag(s) + Br_2(g)}^{\text{oxidation}}$$
$$\underbrace{}_{\text{reduction}}$$

13.4 | Oxidising and Reducing Agents

What is an oxidising agent?

A substance that causes another substance to be oxidised is called an **oxidising agent**. An oxidising agent is reduced when it oxidises another substance.

Oxidising agents	Reducing agents
bromine (Br_2)	carbon (C)
chlorine (Cl_2)	carbon monoxide (CO)
concentrated sulphuric acid (H_2SO_4)	hydrogen (H_2)
nitric acid (HNO_3)	hydrogen sulphide (H_2S)
oxygen (O_2)	metals
potassium manganate(VII) ($KMnO_4$)	potassium iodide (KI)
potassium dichromate(VI) ($K_2Cr_2O_7$)	sulphur dioxide (SO_2)
hydrogen peroxide (H_2O_2)	ammonia (NH_3)

able 13.2 Oxidising and reducing agents

What is a reducing agent?

A substance that causes another substance to be reduced is called a **reducing agent**. The reducing agent is oxidised in the process. Table 13.2 shows some common oxidising and reducing agents.

How can we identify the oxidising and reducing agents in a reaction?

Let us take a look at the reaction between chlorine and hydrogen sulphide. Try and identify the oxidising and reducing agents.

$Cl_2(g) + H_2S(g) \longrightarrow 2HCl(g) + S(s)$

Hydrogen sulphide is oxidised to sulphur. Chlorine is reduced to hydrogen chloride. Chlorine is the oxidising agent because it has oxidised hydrogen sulphide to sulphur. Hydrogen sulphide is the reducing agent because it has reduced chlorine to hydrogen chloride.

oxidised by chlorine

$Cl_2(g) + H_2S(g) \longrightarrow 2HCl(g) + S(s)$

reduced by hydrogen sulphide

How do we identify the oxidising and reducing agents if there is only one product?

The equation for the reaction between hydrogen and oxygen to form water is

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

In a redox reaction, the oxidising agent is reduced and the reducing agent is oxidised. In this reaction, hydrogen has gained oxygen to become water. Therefore it has been oxidised. This means that hydrogen is the reducing agent.

Oxygen has gained hydrogen to become water, so it has been reduced. This means that oxygen is the oxidising agent.

Can oxidising agents and reducing agents be identified based on electron transfer?

Zinc reacts with copper(II) sulphate according to this equation:

zinc + copper(II) sulphate $\longrightarrow$ zinc sulphate + copper

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$$

The ionic equation is shown here.

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

In the reaction, each zinc atom has become a zinc ion, Zn^{2+}.

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$

Each zinc atom has lost two electrons to become a zinc ion, Zn^{2+}. Therefore, zinc has been oxidised. It is the reducing agent. It has reduced the copper(II) ion to copper metal.

We can also examine this reaction in terms of oxidation states. Since the oxidation state of zinc has increased from 0 to +2, we say that zinc has been oxidised and is the reducing agent.

In this same reaction, the copper(II) ion, Cu^{2+}, has accepted electrons and has thus been reduced.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

The oxidation state of copper has decreased from +2 to 0. Thus, we say that copper(II) ion is the oxidising agent.

An oxidising agent removes electrons from another substance. It is reduced during the reaction.

A reducing agent gives up electrons to another substance. It is oxidised during the reaction.

What is the test for an oxidising agent?

Aqueous potassium iodide, KI, is used to test for the presence of an oxidising agent. KI is colourless. If a drop of KI is added to a solution containing an oxidising agent, a brown solution will be formed. The solution turns brown because the iodide ion, I^-, is oxidised to iodine, I_2, by the oxidising agent.

$$\underset{\text{(colourless)}}{2I^-(aq)} \longrightarrow \underset{\text{(brown)}}{I_2(aq)} + 2e^-$$

The iodide ion is colourless, but aqueous iodine, $I_2(aq)$, is brown.

Starch-iodide paper can also be used to test for the presence of oxidising agents. Oxidising agents change the colour of moist starch-iodide paper from white to blue. This is because the iodine produced reacts with the starch to give a blue colour.

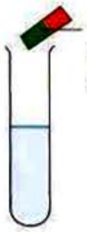


Fig. 13.6 Reaction of potassium dichromate(VI) and sulphur dioxide

What is the test for a reducing agent?

Acidified potassium dichromate(VI) can be used to test for the presence of a reducing agent. Acidified potassium dichromate(VI) is made by adding dilute sulphuric acid to aqueous potassium dichromate(VI). The colour of acidified potassium dichromate(VI) solution changes from orange to green in the presence of a reducing agent. The half equation is shown here.

$$\underset{\text{(orange)}}{Cr_2O_7^{2-}(aq)} + 14H^+(aq) + 6e^- \longrightarrow \underset{\text{(green)}}{2Cr^{3+}(aq)} + 7H_2O(l)$$

In this reaction, the dichromate(VI) ion, $Cr_2O_7^{2-}$, is reduced to the chromium(III) ion, Cr^{3+}. $Cr_2O_7^{2-}$ loses oxygen and the oxidation state of chromium decreases from +6 to +3.

Key ideas

1. A redox reaction is a reaction in which oxidation and reduction occur together.
2. An oxidising agent is a substance that causes oxidation in another substance.
3. A reducing agent is a substance that causes reduction in another substance.
4. Aqueous potassium iodide and starch-iodine paper can be used to test for oxidising agents.
5. Acidified potassium dichromate(VI) can be used to test for reducing agents.

Test Yourself 13.3

Worked Example

A reaction involving hydrogen peroxide, H_2O_2, is shown by the equation below.

$$H_2O_2 + 2H^+ + 2I^- \longrightarrow I_2 + 2H_2O$$

a) Is the reaction with hydrogen peroxide, a redox reaction?
b) Does hydrogen peroxide, H_2O_2, act as an oxidising agent or a reducing agent in this reaction? Explain your answer.

Answer

a) H_2O_2 is reduced. The oxidation state of oxygen is reduced from −1 in H_2O_2 to −2 in H_2O. I^- is oxidised to I_2 because there is an increase in the oxidation state from −1 to 0. Thus, it is a redox reaction.
b) An oxidising agent oxidises another substance and is reduced in the process. H_2O_2 is reduced as it loses oxygen to form water. It is an oxidising agent.

Questions

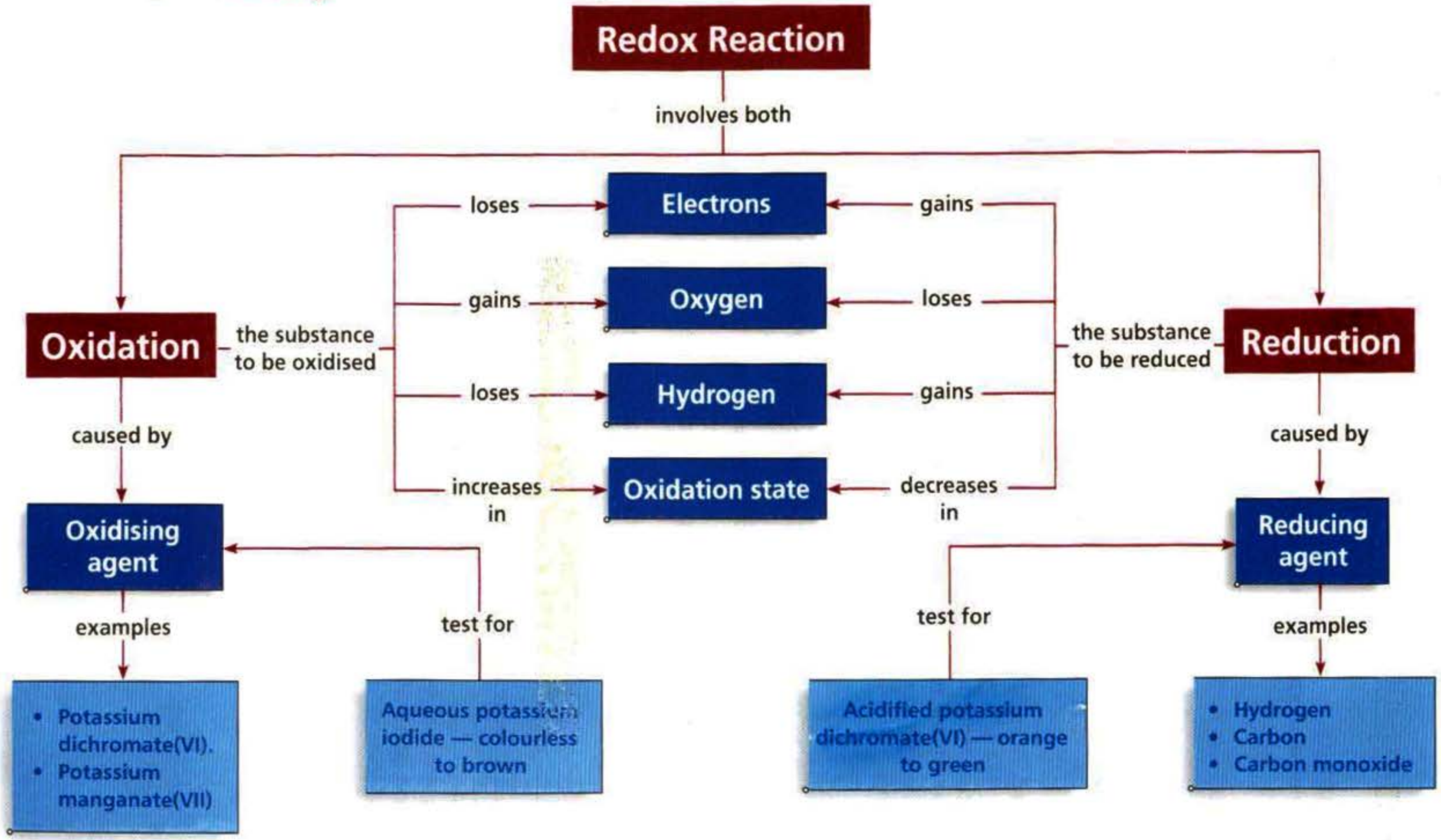
1. Identify the oxidising agent and the reducing agent in each of the following reactions.
   a) bromine + potassium iodide ⟶ potassium bromide + iodine
   b) iron(III) oxide + carbon ⟶ iron + carbon dioxide
   c) ammonia + copper(II) oxide ⟶ copper + water + nitrogen

2. Complete the table comparing the properties of an oxidising agent and a reducing agent.

	Oxidising agent	Reducing agent
Electrons		
Oxidation state		

3. Burning magnesium reacts with carbon dioxide to form a black solid, which is carbon, and a white solid.
   a) Name the white solid.
   b) State with reason, if the reaction is a redox reaction.

Concept Map


Redox Reaction
involves both
Oxidation
the substance to be oxidised
loses
gains
loses
increases in
Electrons
Oxygen
Hydrogen
Oxidation state
gains
loses
gains
decreases in
the substance to be reduced
Reduction
caused by
Oxidising agent
examples
• Potassium dichromate(VI).
• Potassium manganate(VII)
test for
Aqueous potassium iodide — colourless to brown
caused by
Reducing agent
test for
Acidified potassium dichromate(VI) — orange to green
examples
• Hydrogen
• Carbon
• Carbon monoxide

Exercise 13

Foundation

1. In which compound is nitrogen in its lowest oxidation state?

   A N_2O B NH_3
   C NO D NO_2

2. What happens when a bromine atom is changed into a bromide ion?

   A It is oxidised. B It is reduced.
   C It loses a proton. D It loses an electron.

3. In which change has oxidation taken place?

   A $CO_2 \longrightarrow CO_3^{2-}$ B $NO_2 \longrightarrow NO_3^-$
   C $SO_2 \longrightarrow SO_3^{2-}$ D $SO_3 \longrightarrow SO_4^{2-}$

4. Which reaction is **not** an example of reduction?

   A Chlorine to hydrogen chloride.
   B Iron to iron(II) ions.
   C Iron(III) oxide to iron.
   D Nitrogen to ammonia.

5. Which reaction represents a redox reaction?

   A $3O_2 \longrightarrow 2O_3$
   B $Ag^+ + Cl^- \longrightarrow AgCl$
   C $Br_2 + 2I^- \longrightarrow 2Br^- + I_2$
   D $H^+ + OH^- \longrightarrow H_2O$

6. Which statement is true about oxidising agents?

   A Their oxidation state is zero.
   B They are easily oxidised.
   C They never contain hydrogen.
   D They readily accept electrons.

7. Which element is **not** a reducing agent?

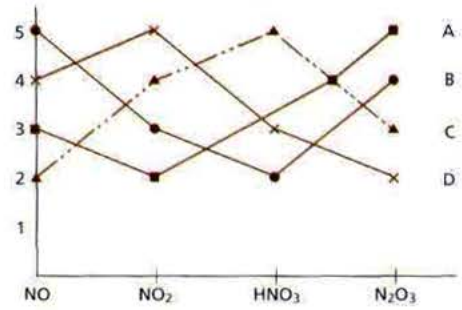
   A Carbon B Chlorine
   C Hydrogen D Magnesium

8. a) Copy and complete the following sentences.
   i) Oxidation is the loss of _____ or the _____ of oxygen.
   ii) Reduction is the gain of _____ or the _____ of oxygen.

   b) State, with reasons, if each of the following reactions are redox reactions.
   i) $2I^-(aq) + Br_2(aq) \longrightarrow I_2(aq) + 2Br^-(aq)$
   ii) $2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$
   iii) $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

Challenge

1. In the changes
   $NO \longrightarrow NO_2 \longrightarrow HNO_3 \longrightarrow N_2O_3$
   which line, A, B, C or D, represents the changes in the oxidation state of nitrogen?



2. When a sample of polluted air was passed through acidified potassium dichromate(VI) the colour of the solution changed from orange to green. Which gas in the air caused this change?

   A Argon B Carbon dioxide
   C Nitrogen D Sulphur dioxide

3. Bromine and sulphur dioxide reacted according to the equation:

   $Br_2 + SO_2 + H_2O \longrightarrow 2HBr + SO_3$

   In the reaction, has the following substance been oxidised or reduced?
   a) Bromine b) Sulphur dioxide

4. Oxidation can be described as the loss of electrons. In the following reactions, has the underlined substance been oxidised or reduced?

   a) $\underline{2Mg(s)} + O_2(g) \longrightarrow 2MgO(s)$
   b) $\underline{2Ag^+(aq)} + Pb(s) \longrightarrow 2Ag(s) + Pb^{2+}(aq)$
   c) $\underline{2NH_3(g)} + 3CuO(s) \longrightarrow 3Cu(s) + N_2(g) + 3H_2O(g)$

5. Which of the reactions are redox reactions?

   a) $CO_2(g) + C(s) \longrightarrow 2CO(g)$
   b) $2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$
   c) $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$
   d) $CuCO_3(s) \longrightarrow CuO(s) + CO_2(g)$
   e) $H_2O(l) + SO_3(g) \longrightarrow H_2SO_4(aq)$
   f) $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
   g) $Ni(s) + Sn^{2+}(aq) \longrightarrow Ni^{2+}(aq) + Sn(s)$

6. What are the colour changes in the following reactions? Complete the table below.

Reaction	Colour change(s)
starch-iodide + reducing agent	
iron(III) chloride to iron(II) chloride	
chlorine to hydrogen chloride	
copper(II) oxide to copper	
potassium dichromate(VI) + oxidising agent	
iron to rust	

7. a) Which of the substances listed will decolourise potassium manganate(VII)?
   i) Hydrogen sulphide
   ii) Iron(II) sulphate
   iii) Nitric acid
   iv) Sulphur dioxide

   b) Bottles of potassium chloride, potassium bromide and potassium iodide were delivered to a chemistry laboratory. However, the bottles were not labelled. Using chlorine or one other chemical, how would you identify the contents of each bottle?

8. a) In the reactions below, has the underlined substance been oxidised, reduced or neither? Explain your answer in terms of transfer of electrons.
   i) $\underline{2Na(s)} + Br_2(g) \longrightarrow 2NaBr(s)$
   ii) $\underline{Pb^{2+}(aq)} + Mg(s) \longrightarrow Mg^{2+}(aq) + Pb(s)$
   iii) $\underline{H^{+}(aq)} + OH^{-}(aq) \longrightarrow H_2O(l)$

   b) When magnesium metal is added to an aqueous solution of iron(III) sulphate, a redox reaction takes place. A grey precipitate of iron and a colourless solution are formed.
   i) Write an ionic equation, including state symbols, for the redox reaction.
   ii) What is the oxidising agent in this reaction? Explain your answer in terms of changes in oxidation states.

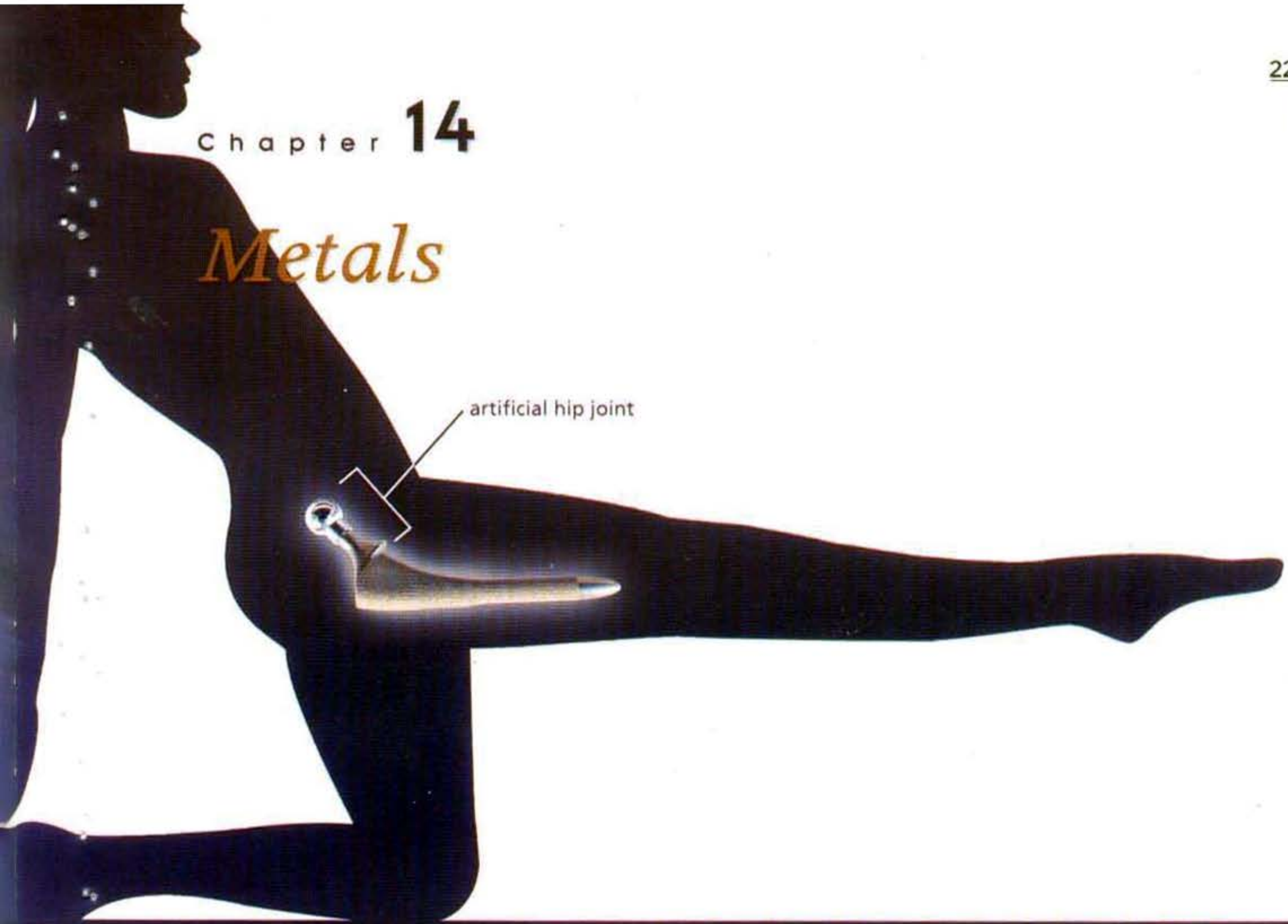
9. Chromium is manufactured by heating chromium(III) oxide (Cr_2O_3) with aluminium.
   a) What does the (III) mean in chromium(III) oxide?
   b) Write the equation for the reaction between chromium(III) oxide and aluminium.
   c) What is the minimum amount (in kg) of chromium(III) oxide needed to make 950 kg of chromium?

10. When copper(I) oxide dissolves in dilute sulphuric acid, copper(II) sulphate and copper are produced.
   a) Write the balanced chemical equation for the reaction.
   b) Discuss the changes in the oxidation states of copper in this reaction.
   c) Is copper(I) oxide oxidised or reduced in this reaction?
   d) This reaction is called a disproportionation reaction. From your answer in (b), deduce one characteristic feature of a disproportionation reaction.

Problem solving strategy:
In the above reaction, the reactant, copper(I) oxide, forms two separate products, copper(II) suphate and copper. Notice the oxidation states of copper in the reactant and the products.

Chapter 14

Metals



The hip is the joint that joins the leg bone to the rest of the skeleton. It leaves the leg free to move. Old people often develop brittle bones and suffer hip fractures when they fall. If the hip is badly damaged, they will be unable to walk. Luckily, advances in medicine and the study of materials now allow doctors to replace a damaged hip with an artificial one made of metal.

An artificial hip must be very strong. It must not wear away easily and must also be light. Artificial joints are commonly made of metal alloys like cobalt-chromium alloys, titanium alloys and stainless steel alloys. Why are pure metals not used instead? You will understand after you learn about the physical and chemical properties of metals in this chapter.

Chapter Outline

14.1 | Metals and Alloys

The physical properties of any substance are determined by the way its particles (atoms, molecules or ions) are packed. Fig. 14.1 shows how atoms are arranged in metals.

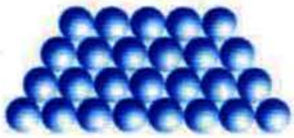


Fig. 14.1 A simple model of the structure of metals

How are the properties of metals related to their structure?

1. Atoms in a metal are packed tightly in layers and are held together by strong **metallic bonds**. Due to these strong metallic bonds, metals usually have *high densities, melting points and boiling points.*

2. In a pure metal, atoms are packed regularly in layers. All these atoms are of the same size. This makes it easy for the layers of atoms to slide over each other when force is applied (Fig. 14.2).

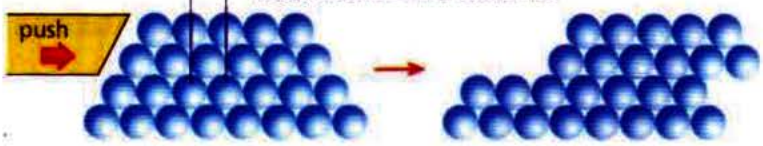


Fig. 14.2 It is easy for the atoms in a pure metal to slide over one another.

This makes metals soft, *ductile* (i.e. they can be drawn into fine wires without breaking) and *malleable* (they can be beaten into thin sheets).

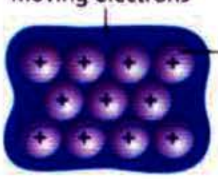


Fig. 14.3 The outermost electrons of the atoms in metals are mobile. Thus, they allow metals to conduct electricity and heat.

3. While the atoms of a metal are tightly packed, the outermost electrons of the atoms can break away easily from the atoms. In other words, the structure of metals can also be described as positive (metal) ions surrounded by a 'sea of mobile electrons'. The mobile electrons allow metals to *conduct electricity* when they are connected to an electrical source. Heat energy is also transferred easily by the mobile electrons in the structure. This makes metals *good conductors of heat.*

Pure metals have many useful properties but they are not widely used. This is because many pure metals

- are soft,
- may react with air and water and wear away easily. We say that these pure metals **corrode** easily (or have a low resistance to **corrosion**).

Most metallic substances used nowadays are **alloys**.

What are alloys?

An **alloy** is *a mixture of a metal with one or a few other elements.* For example,

- bronze is an alloy of copper and tin,
- brass is an alloy of copper and zinc,
- stainless steel is an alloy of iron, chromium, nickel and carbon.

Alloys are made by mixing the molten elements (metals or metals and carbon) in the right proportions and then allowing them to solidify. The alloys produced have more useful physical properties than the pure metals.

Why are metals often used in the form of alloys?

1. Metals can be made harder and stronger by alloying them with other elements. For example, brass is harder and stronger than its constituents, pure copper or pure zinc.

   When a pure metal is alloyed, a different element is added to the pure metal. Atoms of the added element have a different size from those of the pure metal. This breaks up the regular arrangement of atoms in the pure metal. The atoms of different sizes cannot slide over each other easily (Fig. 14.4). This makes the alloy harder and less malleable.

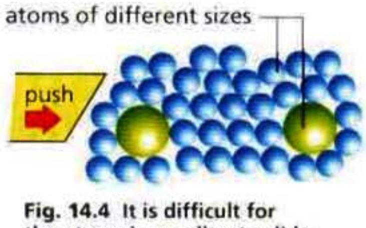


Fig. 14.4 It is difficult for the atoms in an alloy to slide over one another.

2. Alloying can also be used to improve the appearance of the metal. Pewter is an alloy of tin, antimony and copper. It is used to make ornaments and souvenirs because it looks more beautiful than pure tin.

This keychain is made of pewter, an alloy of tin.

3. Alloys are also more resistant to corrosion than pure metals. For example, pure copper corrodes easily. This is why an alloy of copper is used to make coins instead.

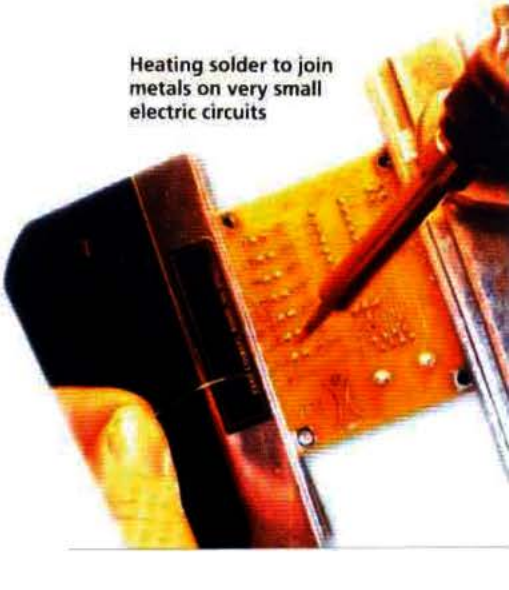
4. Alloying is used to lower the melting points of metals. Solder is an alloy of tin and lead. It has a lower melting point than pure tin or pure lead and can be used to join metals.

Table 14.1 shows examples of alloys, their composition, properties and uses.

Alloy	Composition*	Special properties	Some uses
Brass	copper (70%) zinc (30%)	does not corrode easily, attractive yellow colour like gold	decorative ornaments, musical instruments, coins
Stainless steel	iron (73%) chromium (18%) nickel (8%) carbon (1%)	resistant to corrosion	cutlery, utensils, medical instruments, pipes in chemical industries
Solder	tin (50%) lead (50%)	low melting point	joining metals, e.g. joining metal pipes
Pewter	tin (95%) antimony (3.5%) copper (1.5%)	bright, shiny colour like silver	decorative ornaments

*The percentages of metals used to make each alloy may vary.

Table 14.1 Examples of alloys, their compositions, properties and uses



Heating solder to join metals on very small electric circuits

Copper-nickel alloy is used to make coins. Why is pure copper not used instead?

Key ideas

1. The physical properties of metals are as follows:
   - Usually have high densities, melting points and boiling points
   - Can be bent, stretched or beaten into very thin sheets without breaking
   - Good conductors of heat and electricity
2. An alloy is a mixture of a metal with one or a few other elements.
3. There are four main reasons for making alloys:
   - To improve the strength and hardness of metals
   - To improve the appearance of metals
   - To improve the resistance of metals against corrosion
   - To lower the melting points of metals

Test Yourself 14.1

Worked Example

Which property need not be considered when choosing a metal for making coins?

A Chemical reactivity
B Electrical conductivity
C Hardness
D Melting point

Thought Process

If the metal is too reactive, coins made from it will corrode easily. The chosen metal has to be hard so that the coins do not change shape. It must also have a high melting point so that the coins are not melted down easily. All metals conduct electricity, so this property need not be considered.

Answer

B

Questions

1. Tungsten is used to make the filament inside a bulb because it conducts electricity and has a high melting point. Explain how the structure of metals results in these properties.

2. Draw simple diagrams to show the difference in the arrangement of atoms between copper and brass.

3. 'The properties of alloys are usually different from those of the elements they contain.' Explain why this is so for any two properties.

14.2 | The Reactivity Series

Metals not only have many common physical properties, they also undergo many similar chemical reactions. Some of these reactions are as follows:

1. Form positive ions by the loss of electrons, e.g.
   $Fe \longrightarrow Fe^{2+} + 2e^-$
2. Form ionic compounds, e.g. metal chlorides and metal oxides.
3. Usually react with dilute hydrochloric acid or sulphuric acid to give hydrogen and a salt.
4. React with oxygen to form basic oxides or amphoteric oxides.

However, one metal may react more or less vigorously with a substance than another metal. The metal that reacts more vigorously is said to be more **reactive** than the other metal.

It is useful to arrange metals according to how reactive they are. As such, scientists have worked out a list called the **reactivity series**. In the reactivity series, *metals are listed from the most reactive to the least reactive.*

How is the order of reactivity determined?
Let us determine the reactivity of metals by looking at how they react with water, steam and hydrochloric acid in the laboratory.

The Reaction of Metals with Water

Which metals react with cold water?
Some metals react with cold water to form the metal hydroxide and hydrogen gas.

metal + water ⟶ metal hydroxide + hydrogen

Fig. 14.5 and Fig. 14.6 show what happens when pieces of different metals are dropped into a beaker of cold water.

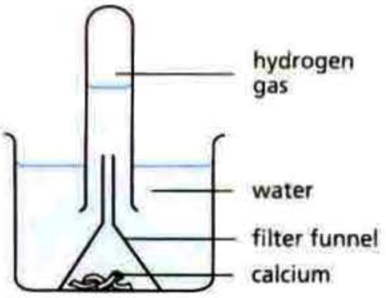


Fig. 14.5 Bubbles of gas are produced. This indicates that calcium reacts readily with cold water.

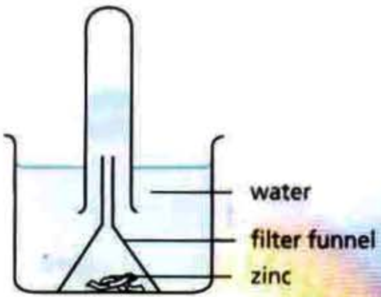


Fig. 14.6 No gas bubbles are produced. Zinc does not react with cold water.

This is a metal part of a ship that sank over a hundred years ago. Why has it worn away?

Table 14.2 gives the observations and chemical equations for the reactions of different metals with cold water.

Metal(s)	Observations	Equation
Potassium	Reacts very violently to form potassium hydroxide and hydrogen gas. Enough heat is produced to ignite the hydrogen gas produced. The hydrogen gas burns with a lilac flame.	$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$
Sodium	Reacts violently to form sodium hydroxide and hydrogen gas. The hydrogen gas formed may catch fire and burn with a yellow flame.	$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$
Calcium	Reacts readily to form calcium hydroxide and hydrogen gas.	$Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$
Magnesium	Reacts very slowly to form magnesium hydroxide and hydrogen gas. A test tube of hydrogen gas is produced only after a few days.	$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(s) + H_2(g)$
Zinc Iron* Lead Copper Silver	No reaction occurs.	

*Iron reacts with water very slowly in the presence of air, in a process called rusting. See section 14.5.

Table 14.2 Reactions of metals with cold water

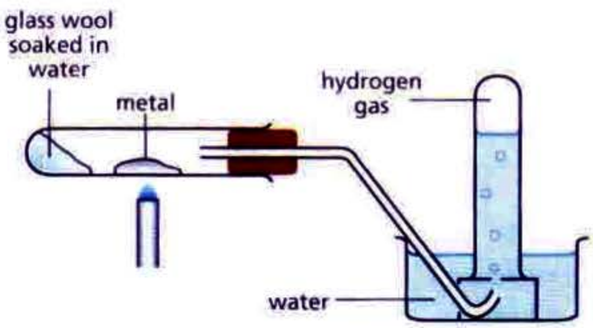
As you can see from how vigorous the reactions are, the most reactive metals are potassium, followed by sodium and then calcium.

Link

If metals such as potassium, sodium and calcium react with cold water, we can expect them to have even more violent reactions with hot water or steam. Find out why a chemical reaction is affected by temperature in chapter 18.

Which metals only react with steam?

Zinc and iron do not react with cold water but they do react with steam. Fig. 14.7 shows the apparatus used for reacting a metal with steam.



To prevent water from being sucked back into the hot test tube, the delivery tube must be removed from the water in the trough before turning off the Bunsen flame.

Fig. 14.7 Reacting a metal with steam

The metal is strongly heated until it is very hot. The glass wool soaked in water is also heated to generate a flow of steam over the hot metal.

Metals react with steam to form the metal oxide and hydrogen gas.

metal + steam ⟶ metal oxide + hydrogen

Table 14.3 shows the observations and chemical equations for the reactions of some metals with steam.

Metal(s)	Observations	Equation
Magnesium	Hot magnesium reacts violently with steam to form magnesium oxide (a white powder) and hydrogen gas. A bright white glow is produced during the reaction.	$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$
Zinc	Hot zinc reacts readily with steam to produce zinc oxide and hydrogen gas. Zinc oxide is yellow when hot and white when cold.	$Zn(s) + H_2O(g) \longrightarrow ZnO(s) + H_2(g)$
Iron	Red-hot iron reacts slowly with steam to form iron oxide and hydrogen gas. The iron must be heated constantly in order for the reaction to proceed.	$3Fe(s) + 4H_2O(g) \longrightarrow Fe_3O_4(s) + 4H_2(g)$
Lead Copper Silver	No reaction occurs.	

Table 14.3 Reactions of some metals with steam

From the experimental results shown in Tables 14.2 and 14.3, we can conclude that
- potassium, sodium and calcium are highly reactive metals.
- magnesium, zinc and iron are fairly reactive metals.
- lead, copper and silver are unreactive metals.
- reactive metals tend to react to form compounds, while unreactive metals tend to remain as metals.
- the reactivity of metals is in the order shown in Fig. 14.8.

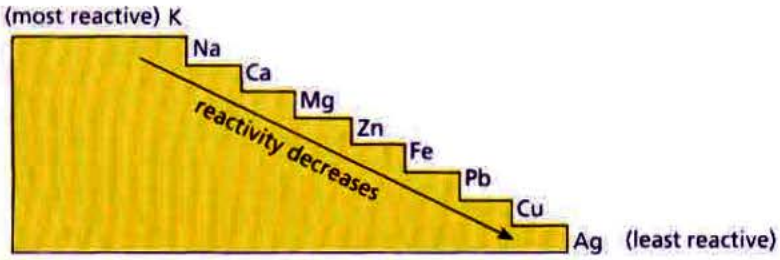


Fig. 14.8 Order of reactivity of metals

Quick Check

Explain why coins made of zinc look dull after some time but silver coins remain bright and shiny for years.

Link

Do metals react with dilute sulphuric acid in the same way they react with dilute hydrochloric acid? Recall what you have learnt in chapter 11.

The Reaction of Metals with Dilute Hydrochloric Acid

Many metals react with dilute acids to produce a salt and hydrogen.

metal + dilute acid $\longrightarrow$ salt + hydrogen

When a metal reacts with dilute hydrochloric acid, the products are the metal chloride and hydrogen gas.

metal + dilute hydrochloric acid $\longrightarrow$ metal chloride + hydrogen

The reactions of different metals with dilute hydrochloric acid also indicate how reactive the metals are.

Table 14.4 summarises the reactions of metals with dilute hydrochloric acid.

Metal(s)	Observations	Equation
Potassium Sodium	Explosive reaction. These reactions should not be carried out in the school laboratory.	$2K(s) + 2HCl(aq) \longrightarrow 2KCl(aq) + H_2(g)$ $2Na(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2(g)$
Calcium	Reacts violently to give hydrogen gas.	$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$
Magnesium	Reacts rapidly to give hydrogen gas.	$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
Zinc	Reacts moderately fast to give hydrogen gas.	$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$
Iron	Reacts slowly to give hydrogen gas.	$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$
Lead Copper Silver	No reaction occurs.	

Table 14.4 Reactions of metals with dilute hydrochloric acid

The experimental results shown in Table 14.4 confirm that the reactivity of metals is in the order shown by the reactivity series below (Fig. 14.9).

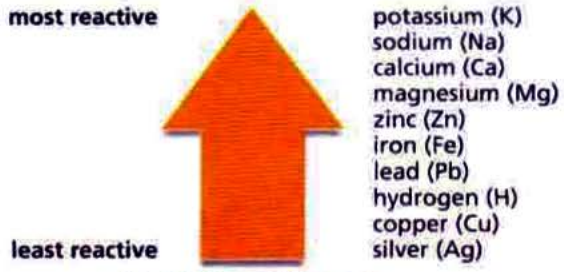


Fig.14.9 Reactivity series

The reactivity series also includes non-metals such as hydrogen. Hydrogen is placed between lead and copper. Only metals that are more reactive than hydrogen react with dilute acids to produce hydrogen gas.

The Reduction of Metal Oxides with Carbon

We have compared the reactivity of metals by studying their reactions with water and hydrochloric acid. We can also compare the reactivity of different metals by studying how easily metal oxides decompose. The more reactive a metal is, the more difficult it is to decompose its oxides — reduce the oxide to the metal.

When a mixture of copper(II) oxide and carbon is heated (Fig. 14.10), copper(II) oxide is reduced to copper and carbon is oxidised to carbon dioxide. The equation for the reaction is

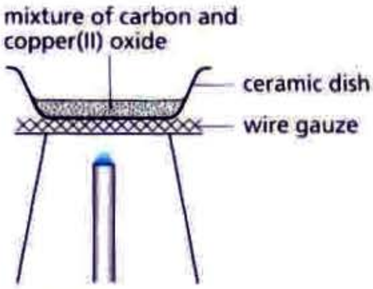


Fig. 14.10 Apparatus used to reduce copper(II) oxide with carbon

$$\underset{\text{(black solid)}}{\text{copper(II) oxide}} + \text{carbon} \xrightarrow{\text{heat}} \underset{\text{(reddish-brown solid)}}{\text{copper}} + \text{carbon dioxide}$$

$$2CuO(s) + C(s) \xrightarrow{\text{heat}} 2Cu(s) + CO_2(g)$$

If the experiment is repeated with magnesium oxide, no reaction takes place. The reactions of some metal oxides with carbon are shown in Table 14.5.

Metal oxide	Reaction
Potassium oxide (K_2O) Sodium oxide (Na_2O) Calcium oxide (CaO) Magnesium oxide (MgO)	Oxides are not reduced by carbon.
Zinc oxide (ZnO) Iron(II) oxide (FeO) Lead(II) oxide (PbO) Copper(II) oxide (CuO)	Oxides are reduced by carbon.
Silver oxide (Ag_2O)	Oxide is reduced by heating.

Table 14.5 Reduction of metal oxides with carbon

If a metal is below copper in the reactivity series, its oxide will decompose simply by heating, without the need for a reducing agent such as carbon. For example, silver oxide decomposes into silver and oxygen when it is heated to 200 °C as shown in this equation:

$$\text{silver oxide} \xrightarrow{\text{heat}} \text{silver} + \text{oxygen}$$

$$2Ag_2O(s) \xrightarrow{\text{heat}} 4Ag(s) + O_2(g)$$

What is the importance of this reaction in industry?

Metals need to be extracted from their ores, which are compounds found naturally. These ores are mainly made up of metal oxides. In industry, metals that are below magnesium in the reactivity series are often extracted from their ores by reduction with carbon.

Zinc and lead are extracted by reduction with carbon, as represented by the equations below:

zinc oxide + carbon ⟶ zinc + carbon monoxide

$ZnO(s) + C(s) \longrightarrow Zn(s) + CO(g)$

lead(II) oxide + carbon ⟶ lead + carbon monoxide

$PbO(s) + C(s) \longrightarrow Pb(s) + CO(g)$

Iron is also extracted from its ore by reduction with carbon. We shall learn the details of this process in section 14.3.

The oxides of metals that are above zinc in the reactivity series are not reduced by carbon. These oxides are so stable that they can only be reduced by passing electricity through them.

The Reduction of Metal Oxides with Hydrogen

Besides carbon, hydrogen can also be used for reducing metal oxides to metals. Fig. 14.11 shows the apparatus used for the reduction of a metal oxide to the metal using hydrogen.

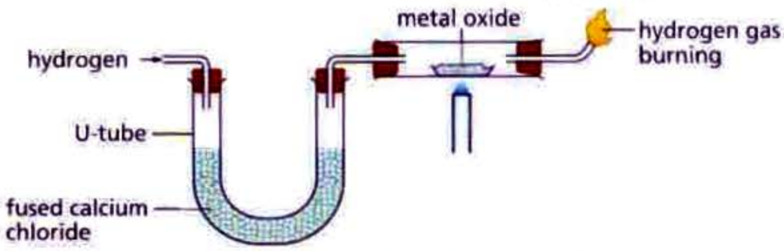


Fig. 14.11 The reduction of a metal oxide with a stream of dry hydrogen gas

Hydrogen gas is passed over the metal oxide. It acts as a reducing agent. Hydrogen reduces oxides of some metals, such as iron, copper and silver, to the metals as shown by the equation:

metal oxide + hydrogen ⟶ metal + steam

Oxides of reactive metals such as potassium, sodium, calcium, magnesium and zinc, are not reduced by hydrogen. Table 14.6 summarises the reactions of metal oxides with hydrogen.

Metal oxide	Reaction with hydrogen
Potassium oxide (K_2O) Sodium oxide (Na_2O) Calcium oxide (CaO) Magnesium oxide (MgO) Zinc oxide (ZnO)	Heated metal oxides are not reduced.
Iron(II) oxide (FeO) Lead(II) oxide (PbO) Copper(II) oxide (CuO) Silver oxide (Ag_2O)	Heated metal oxides are reduced.

Table 14.6 The reduction of metal oxides with hydrogen

Key ideas

1. A summary of the reactions of metals with water and dilute acids:

Metal(s)	Reaction with water or steam	Reaction with dilute acid
potassium sodium	react with cold water and steam	explosive reaction
calcium	reacts with cold water and steam	violent reaction
magnesium zinc	react with steam	moderately fast reaction
iron	reacts with steam	slow reaction
lead copper silver	no reaction with water or steam	no reaction

2. Metals are listed in the reactivity series from the most reactive to the least reactive.

Reactivity series

most reactive
K
Na
Ca
Mg
Zn
Fe
Pb
(H)
Cu
Ag
least reactive

3. Oxides of metals above zinc cannot be reduced by heating with carbon.
4. Oxides of metals above iron cannot be reduced by heating with hydrogen.

Test Yourself 14.2

Worked Example

Aluminium (Al) is above zinc but below magnesium in the reactivity series. Predict the reaction between aluminium and hydrochloric acid.

Thought Process

Order of reactivity: magnesium > aluminium > zinc. Since magnesium and zinc react rapidly with hydrochloric acid to form their metal chlorides and hydrogen, aluminium will undergo the same reaction.

Answer

Aluminium will react rapidly with hydrochloric acid to form aluminium chloride ($AlCl_3$) and hydrogen.

Questions

1. Manganese (Mn) is between magnesium and iron in the reactivity series. It reacts with steam to form manganese(II) oxide. Write the equation for the reaction.

2. Why are hot water tanks made of copper and not steel?

3. a) X is a metal which does not react with cold water but reacts with steam to form a white oxide. Which **two** metals could be X?
   b) The reaction between X and dilute hydrochloric acid is moderately fast at room temperature. From this observation, deduce the identity of X.

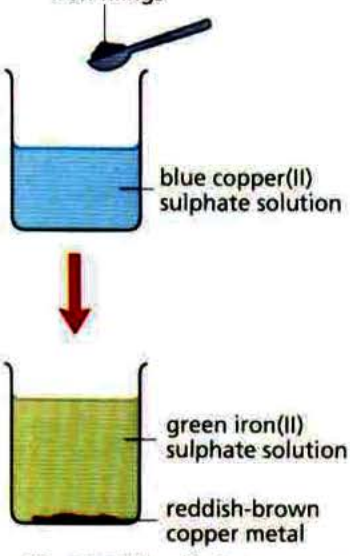


Fig. 14.12 Iron displaces copper from copper(II) sulphate solution.

The Displacement Reactions of Metals

More reactive metals can displace less reactive metals from their salt solutions. For example, solid iron displaces copper ions from a solution of copper(II) sulphate (Fig. 14.12). Due to the action of iron, copper metal is precipitated out of the solution as a pink or a reddish-brown solid.

The reaction can be represented by the equation below:

iron + copper(II) sulphate $\longrightarrow$ iron(II) sulphate + copper

$$Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$$

The ionic equation is

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$$

In other words, atoms of the more reactive metal become ions and form compounds while ions of the less reactive metal change back to atoms. The following experiment investigates more of these displacement reactions.

Chem-Aid

1. Copper often appears pink when it is freshly formed (displaced).
2. Iron(II) sulphate is prepared using excess acid so that it does not break down easily.

Experiment 1

To investigate the displacement of metals from their solutions by another metal.

Procedure

1. Solutions of the following salts are prepared in separate test tubes.
   - Copper(II) sulphate
   - Lead(II) nitrate
   - Zinc sulphate
   - Iron(II) sulphate
   - Magnesium sulphate
2. A clean strip of copper is placed into each solution. What happens then is recorded.
3. The experiment is repeated with clean strips of lead, zinc, magnesium and iron nails using fresh salt solutions.

Table 14.7 shows some of the observations for the experiment.

Salt solution / Metal	Magnesium sulphate	Zinc sulphate	Copper(II) sulphate	Lead(II) nitrate	Iron(II) sulphate
Magnesium		Solution remains colourless. Grey deposit of zinc formed on magnesium. **$Mg(s)$ + $ZnSO_4(aq)$ ⟶ $Zn(s)$ + $MgSO_4(aq)$**	Blue solution turns colourless. Reddish-brown deposit of copper formed on magnesium. **$Mg(s)$ + $CuSO_4(aq)$ ⟶ $Cu(s)$ + $MgSO_4(aq)$**	Solution remains colourless. Grey deposit of lead formed on magnesium. **$Mg(s)$ + $Pb(NO_3)_2(aq)$ ⟶ $Pb(s)$ + $Mg(NO_3)_2(aq)$**	Pale green solution turns colourless. Grey deposit of iron formed on magnesium. **$Mg(s)$ + $FeSO_4(aq)$ ⟶ $Fe(s)$ + $MgSO_4(aq)$**
Zinc	No reaction.		Blue solution turns colourless. Reddish-brown deposit of copper formed on zinc. **$Zn(s)$ + $CuSO_4(aq)$ ⟶ $Cu(s)$ + $ZnSO_4(aq)$**	Solution remains colourless. Grey deposit of lead formed on zinc. **$Zn(s)$ + $Pb(NO_3)_2(aq)$ ⟶ $Pb(s)$ + $Zn(NO_3)_2(aq)$**	Pale green solution turns colourless. Grey deposit of iron formed on zinc. **$Zn(s)$ + $FeSO_4(aq)$ ⟶ $Fe(s)$ + $ZnSO_4(aq)$**
Copper	No reaction.	No reaction.		No reaction.	No reaction.
Lead	No reaction.	No reaction.	Blue solution turns colourless. Reddish-brown deposit of copper formed on lead. **$Pb(s)$ + $CuSO_4(aq)$ ⟶ $Cu(s)$ + $PbSO_4(aq)$**		No reaction
Iron	No reaction.	No reaction.	Blue solution turns pale green. Reddish-brown deposit of copper formed on iron. **$Fe(s)$ + $CuSO_4(aq)$ ⟶ $Cu(s)$ + $FeSO_4(aq)$**	Colourless solution turns pale green. Grey deposit of lead formed on iron. **$Fe(s)$ + $Pb(NO_3)_2(aq)$ ⟶ $Pb(s)$ + $Fe(NO_3)_2(aq)$**	

Table 14.7 Observations for some displacement reactions

From the results in Table 14.7, you can see that a *metal higher up in the reactivity series will displace a metal that is lower in the series from its salt solution.*

Metal displacement reactions are redox reactions

In a metal displacement reaction, the more reactive metal is oxidised while the less reactive metal is reduced. Thus, a displacement reaction is a **redox reaction**.

Link

Displacement reactions also take place between non-metals, for example, between the halogens in Group VII of the Periodic Table. Find out more in chapter 16.

For example, in the reaction between iron and copper(II) sulphate, the more reactive metal, iron, reduces the copper(II) ions to copper. At the same time, iron itself is oxidised to the iron(II) ions. The ionic equation for the reaction is

Copper(II) ions are reduced to become copper atoms.

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$$

Iron atoms are oxidised to become iron(II) ions.

A similar redox reaction occurs when a piece of magnesium is dipped into a copper(II) salt solution to displace copper. The more reactive metal, magnesium, reduces the copper(II) ions to copper. Magnesium itself is oxidised to the magnesium ions.

Copper(II) ions are reduced to become copper atoms.

$$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$$

Magnesium atoms are oxidised to become magnesium ions.

On the other hand, if a piece of copper foil is dipped into magnesium sulphate solution, no reaction occurs.

no change

$$Cu(s) + Mg^{2+}(aq) \longrightarrow Cu(s) + Mg^{2+}(aq)$$

no change

You can see from the above ionic equations that a metal higher up in the reactivity series has a greater tendency to form positive ions. This is why a more reactive metal can displace a less reactive metal from its salt solution. Copper is less reactive than magnesium. It has a lower tendency to form ions compared to magnesium. Thus, it cannot displace magnesium from its salt solution.

Reaction Between a Metal and the Oxide of Another Metal

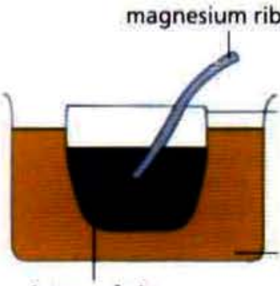


Fig. 14.13 Reaction between zinc and copper(II) oxide

A more reactive metal has a higher tendency to form its positive ions compared to a less reactive metal. This is why a more reactive metal can reduce the oxide of a less reactive metal. An example is the reaction between zinc and copper(II) oxide. Fig. 14.13 shows the apparatus used to study the reaction.

The magnesium ribbon acts as a fuse. When the magnesium ribbon is ignited, it provides enough energy to start the reaction between zinc and copper(II) oxide to form zinc oxide and copper.

Since zinc is more reactive than copper, it will form positive ions more readily. Zn^{2+} ions react with O^{2-} ions from copper(II) oxide to form zinc oxide. In the process, copper(II) oxide is reduced to copper.

zinc + copper(II) oxide ⟶ zinc oxide + copper

$$Zn(s) + CuO(s) \longrightarrow ZnO(s) + Cu(s)$$

The more reactive the metal is, the more readily it forms compounds. On the other hand, unreactive metals tend to stay uncombined. Therefore, if the experiment shown in Fig. 14.13 is repeated using a mixture of zinc oxide and copper powder, no reaction will occur.

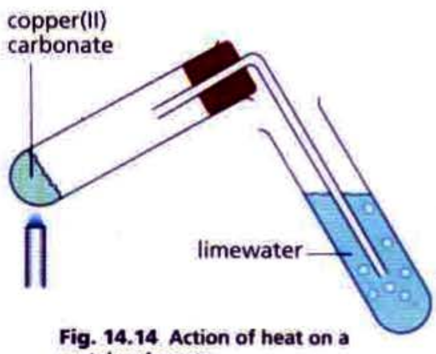


Fig. 14.14 Action of heat on a metal carbonate

Action of Heat on Metal Carbonates

Some compounds are more difficult to decompose by heat than others. This means that these compounds are more stable to heat than others. The thermal stability of metal carbonates can be tested by heating them in a dry test tube (Fig. 14.14). Table 14.8 summarises the action of heat on some metal carbonates.

Metal carbonate	Observation
Potassium carbonate, K_2CO_3 Sodium carbonate, Na_2CO_3	unaffected by heat
Calcium carbonate, $CaCO_3$ Magnesium carbonate, $MgCO_3$ Zinc carbonate, $ZnCO_3$ Iron(II) carbonate, $FeCO_3$ Lead(II) carbonate, $PbCO_3$ Copper(II) carbonate, $CuCO_3$	decompose into metal oxide and carbon dioxide on heating
Silver carbonate, Ag_2CO_3	decomposes into silver and carbon dioxide on heating

Table 14.8 Action of heat on metal carbonates

You can see that the thermal stability of the metal carbonates is related to the position of the metal in the reactivity series. *The more reactive the metal is, the more difficult it is to decompose its compounds.* Therefore, sodium and potassium carbonates are not affected by heat.

The carbonates of metals below sodium in the reactivity series decompose to form the oxides of the metals and carbon dioxide. In the case of silver carbonate, the silver oxide produced is thermally unstable. It further decomposes to form silver.

Using the Reactivity Series

The reactivity series is useful for

1. predicting the behaviour of a metal from its position in the reactivity series,
2. predicting the position of an unfamiliar metal in the reactivity series from a given set of experimental results.

Example 1

Tin is below iron but above lead in the reactivity series. Predict the reaction between tin(II) oxide, SnO, and

a) carbon.
b) magnesium.

Solution:

a) Order of reactivity: iron > tin > lead. Since the oxides of both iron and lead can be reduced by carbon, tin(II) oxide will undergo the same reaction. Tin(II) oxide will be reduced by carbon on heating to form tin and carbon monoxide.

b) Magnesium is more reactive than tin. It has a higher tendency to form its positive ions (and compounds). Thus, it will reduce tin(II) oxide to tin and form magnesium oxide at the same time. Tin(II) oxide will react with magnesium to form magnesium oxide and tin.

Example 2

When a piece of chromium is placed in zinc sulphate solution, no reaction occurs. When a mixture of chromium powder and iron(III) oxide is heated strongly, a reaction takes place. Deduce the position of chromium in the reactivity series.

Solution:

Since chromium does not react with zinc sulphate solution, it must be less reactive than zinc, i.e. below zinc in the reactivity series. Since it reacts with iron(III) oxide, it must be more reactive than iron. Thus, chromium is below zinc but above iron in the reactivity series.

Knowing the reactivity series also helps us to choose suitable methods for extracting metals and to prevent iron from rusting.

Key ideas

A metal that is higher up in the reactivity series has a greater tendency to form positive ions than a metal lower in the series.

- A metal that is higher up in the reactivity series will displace a metal that is lower in the series from the salt solution or oxide of the less reactive metal.
- A more reactive metal will form compounds more readily than a less reactive metal.
- Carbonates of metals higher up in the reactivity series are more thermally stable than carbonates of metals lower in the reactivity series.

Test Yourself 14.3

Questions

1. The following experiments were carried out on the metals *W*, *X*, *Y* and *Z*.
   - *Z* displaces *W* from the salt solution of the metal *W*.
   - *W* displaces *X* from the salt solution of the metal *X*.
   - *X* displaces *Y* from the salt solution of the metal *Y*.

   What is the order of reactivity of these metals, placing the most reactive metal first?

   A $W > X > Y > Z$
   B $X > W > Y > Z$
   C $Y > W > X > Z$
   D $Z > W > X > Y$

2. Describe the colour change in copper(II) carbonate when it is heated strongly. What happens when dilute sulphuric acid is added to the heated salt? Write balanced equations for the reactions.

14.3 | Extracting Metals

Only a small number of metals (the very unreactive ones), such as gold and platinum, occur freely in nature as uncombined metals. Most metals react with other elements to form **ores**. Obtaining metals from their ores generally involves three major stages:

1. Concentrating the ore
2. Extracting crude metal from the ore
3. Refining crude metal to obtain the pure metal (see chapter 15)

Fig. 14.15 The stages involved in obtaining metals from their ores

Concentrating metal ores

An **ore** is *a compound of the metal* (usually the oxides, sulphides, chlorides or carbonates) *mixed with large amounts of earth and rock*. Earth and rock are removed before metal is extracted from the ore. This results in a metal ore that contains little waste material.

Extracting metal from ores

There are two main methods for extracting metals from their ores:

1. Reducing the metal compound (ore) to the metal using carbon
2. Using electricity to decompose the molten metal compound (ore) to the metal

The position of a metal in the reactivity series determines the method used for its extraction. The methods used for extracting some common metals are summarised below.

Chem-Aid

Non-metals such as hydrogen and carbon are also listed in the reactivity series. Carbon is placed above zinc and below aluminium. Therefore, carbon can reduce the compounds of zinc and compounds of metals below zinc in the reactivity series.

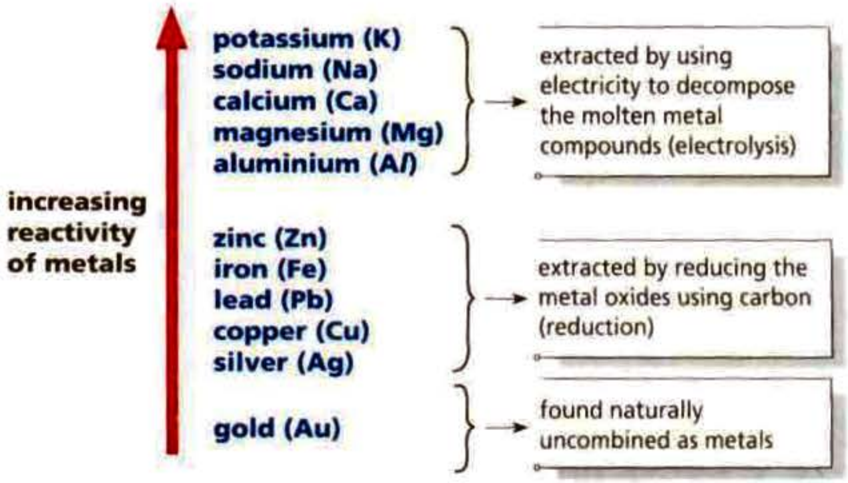


Fig. 14.16 Methods of extracting metals based on their reactivity

Try it Out

The five most common metals in the Earth's crust are aluminium, iron, magnesium, manganese and titanium. Find out the names of the ores from which magnesium, manganese and titanium are extracted.

In general, the more reactive the metal is, the harder it is to extract the metal from its ore. Reactive metals such as sodium, potassium, calcium, magnesium and aluminium cannot be extracted by reduction with carbon. The compounds of these metals are very difficult to split up. These metals are extracted using electricity, in a process called electrolysis (see chapter 15).

The metals placed in the middle of the reactivity series, such as zinc and iron, are not so reactive. They are readily extracted by reducing their oxides with carbon. The metals lowest in the reactivity series, such as gold, can be found in nature as uncombined elements.

Extracting Iron from Haematile

The main ore of iron is **haematite**. Haematite contains iron(III) oxide mixed with impurities such as sand and clay. Iron is extracted from haematite in a **blast furnace** (Fig. 14.17). Haematite, coke (which is mainly carbon) and limestone (calcium carbonate) are added at the top of the blast furnace. Blasts of hot air are blown into the furnace near the bottom.

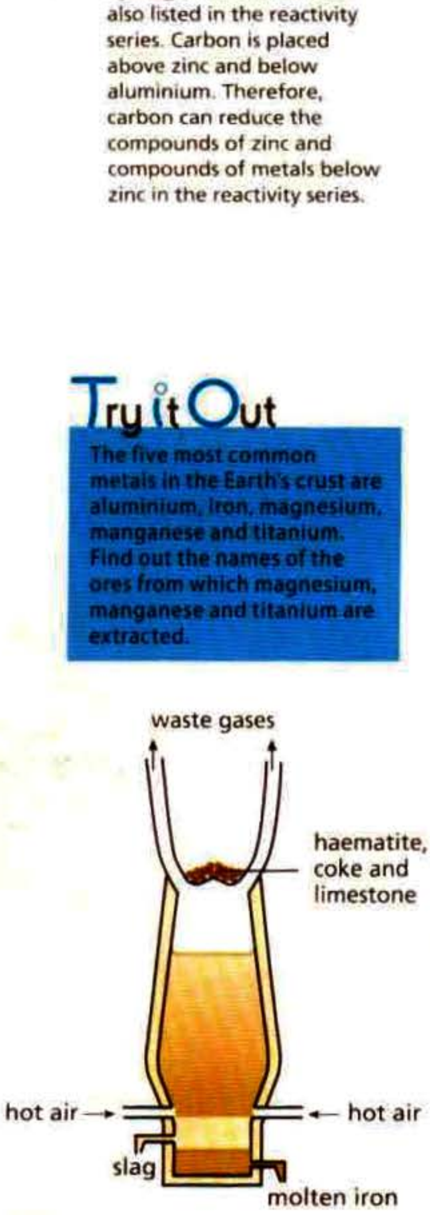


Fig. 14.17 A simple diagram of a blast furnace

Fig. 14.18 is a simple flow chart of what goes on in a blast furnace.

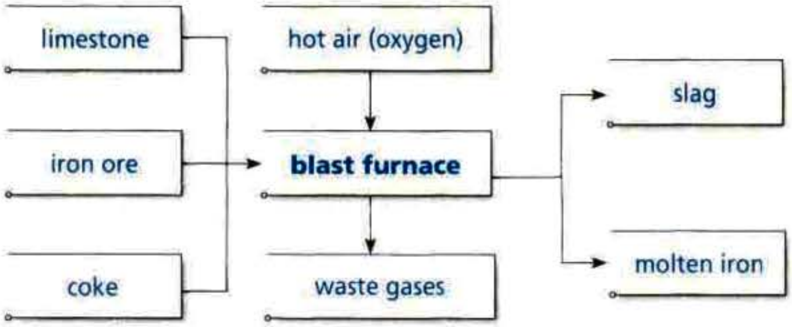


Fig. 14.18 A simple flow chart for extracting iron in a blast furnace

What chemical reactions take place in the blast furnace?
In the blast furnace, a series of chemical reactions takes place.

1. **Carbon dioxide is produced.**
The carbon in coke burns in a blast of hot air to produce carbon dioxide. This reaction produces a lot of heat.

carbon + oxygen $\xrightarrow{\text{heat}}$ carbon dioxide

$C(s) + O_2(g) \xrightarrow{\text{heat}} CO_2(g)$

The limestone (calcium carbonate) is decomposed by heat to produce carbon dioxide and calcium oxide (quicklime).

calcium carbonate $\xrightarrow{\text{heat}}$ calcium oxide + carbon dioxide

$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$

2. **Carbon monoxide is produced.**
As the carbon dioxide rises up the furnace, it reacts with more coke to form carbon monoxide.

carbon + carbon dioxide $\xrightarrow{\text{heat}}$ carbon monoxide

$C(s) + CO_2(g) \xrightarrow{\text{heat}} 2CO(g)$

3. **Haematite is reduced to iron.**
The carbon monoxide reduces the iron(III) oxide in haematite to iron.

iron(III) oxide + carbon monoxide $\xrightarrow{\text{heat}}$ molten iron + carbon dioxide

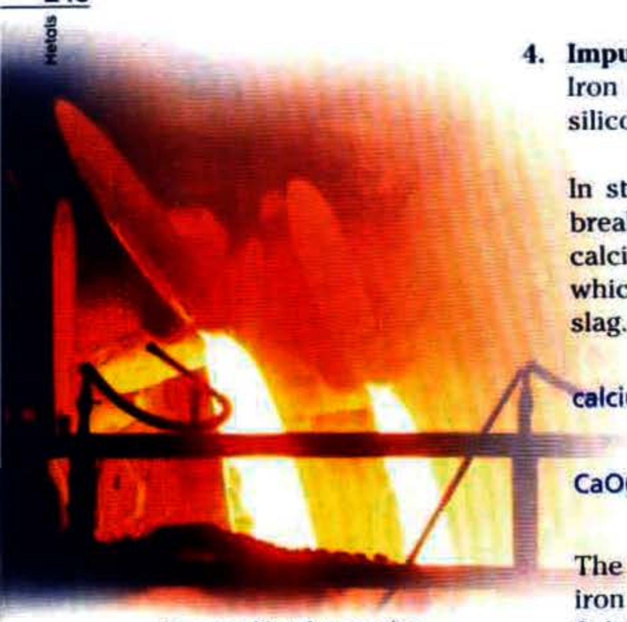
$Fe_2O_3(s) + 3CO(g) \xrightarrow{\text{heat}} 2Fe(l) + 3CO_2(g)$

The iron formed is molten and runs to the bottom of the furnace. Hot waste gases containing carbon monoxide, carbon dioxide and nitrogen escape through the top of the furnace.

Chem-Aid

Upon heating, calcium carbonate breaks down into calcium oxide and carbon dioxide. This process is called thermal decomposition.

A blast furnace can be 50 – 70 m high. It is constructed from steel and lined with fireproof bricks inside.



Dumping blast furnace slag

4. Impurities are removed.

Iron ore contains the impurities sand and clay, which are silicon oxides. Limestone is added to remove these impurities.

In step 1, you saw that the limestone (calcium carbonate) breaks down to form calcium oxide and carbon dioxide. The calcium oxide, a basic oxide, reacts with silicon(IV) oxide which is acidic, and with other impurities, to form a molten slag.

$$\text{calcium oxide} + \text{silicon(IV) oxide} \xrightarrow{\text{heat}} \underset{\text{(slag)}}{\text{calcium silicate}}$$

$$CaO(s) + SiO_2(s) \xrightarrow{\text{heat}} CaSiO_3(l)$$

The lighter slag floats on top of the molten iron. The slag and iron are tapped off separately at the bottom of the furnace. Solidified slag is mainly used for road surfacing.

Key ideas

1. The method used to extract a metal from its ore depends on the position of the metal in the reactivity series.
   - Metals higher up in the series need to be extracted using electricity.
   - Metals lower in the series can be extracted by reduction with carbon.
2. Iron is extracted from haematite (Fe_2O_3) in the blast furnace by reduction with coke (carbon). Limestone (calcium carbonate) is added to remove impurities.

Test Yourself 14.4

Worked Example

Which substance is a reducing agent in the blast furnace for the production of iron?

A Air
B Calcium carbonate
C Carbon monoxide
D Silicon(IV) oxide

Thought Process

A reducing agent removes oxygen from another substance. Carbon monoxide removes oxygen from iron(III) oxide (in haematite) to form iron.

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

Answer

C

Questions

1. The positions of barium and tin in the reactivity series are shown below.

barium
calcium
iron
tin
copper

↑ **increasing reactivity**

Suggest the method that would be most suitable for extracting these metals from their respective ores
a) Barium
b) Tin

2. Name the following substances present in the blast furnace during the extraction of iron from haematite.
a) **Four** ionic compounds
b) **Two** reducing agents

14.4 | The Uses of Iron and Steel

The iron that is extracted from the blast furnace is known as cast iron (or pig iron). Most of it goes into producing steel.

What is steel?
Steel is an alloy of iron with carbon and/or other metals. There are many types of steel. Depending on the amount of carbon and other metals added, and the type of metals added to iron, each type of steel has its own special properties and uses.

How is steel made from cast iron?
Steel-making involves oxidation, followed by alloying. The impurities in iron are removed by passing pure oxygen through molten cast iron. Carbon and small amounts of other metals are then added to produce different types of steel with special properties (Fig. 14.19).

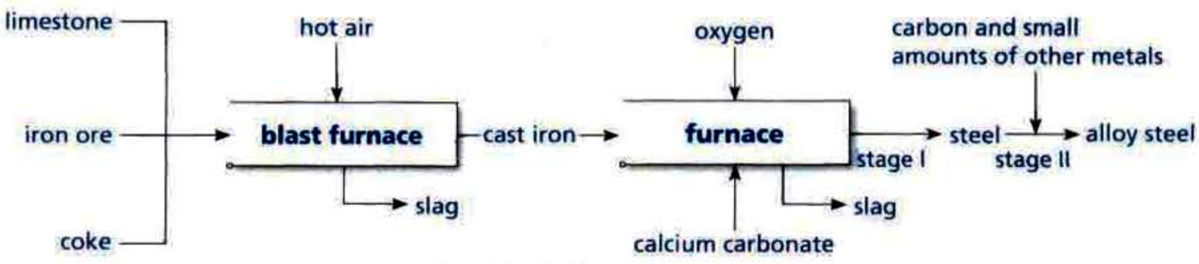


Fig. 14.19 Making steel from iron ore

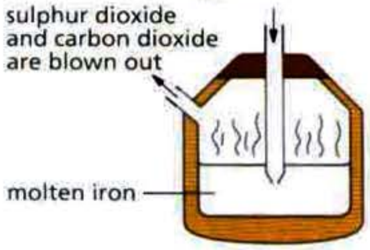


Fig. 14.20 Purifying cast iron

Quick Check

Most metals can be recycled. State three reasons why the recycling of metals is important in modern society.

Car bodies are made of mild steel.

Stainless steel is used to make surgical tools because it does not rust.

1. **Removing impurities by oxidation**
   The common impurities in cast iron are non-metals such as carbon, sulphur, phosphorus and silicon. To remove these impurities, cast iron is first melted and poured into a large container (Fig. 14.20). Pure oxygen at high pressure is then blown through the molten iron to oxidise the impurities. The impurities form gases, such as sulphur dioxide and carbon dioxide, that are blown out of the furnace.

   Calcium carbonate is later added to the furnace. At high temperatures, calcium carbonate decomposes to form calcium oxide which reacts with silicon(IV) oxide and phosphorus oxide to form slag.

   The slag can be poured off. The molten metal left behind is pure iron, also known as wrought iron. Wrought iron is soft and bends easily. It is used for making garden gates and chains.

2. **Adding other elements to make the various types of steel**
   Carbon and other metals are now stirred into the molten iron in the correct amounts to make steel with special properties.

What are the uses of steel?
Although there are many different types of steel, it is convenient to group them into two main categories — carbon steels and alloy steels.

Carbon steels
These are general purpose steels which contain mainly iron and carbon. Mild steel (low carbon steel) contains 99.5% iron and up to 0.25% carbon. It is strong and malleable. It is used to make car bodies and machinery.

High carbon steel contains 0.45 – 1.5% carbon. It is strong but brittle. High carbon steel is used to make knives, hammers, chisels, saws and other cutting and boring tools.

Alloy steels
Alloy steels consist of iron and carbon together with one or more of the following metals: manganese (Mn), nickel (Ni), chromium (Cr), tungsten (W) and vanadium (V). The reason for adding these metals is to change the properties of the steel.

For example, in manganese steel used for making springs and drills, manganese is added to increase its strength and hardness. Nickel and chromium are added to increase resistance to corrosion.

Stainless steel is an alloy of iron, chromium, nickel and a little carbon. It does not rust and is, therefore, used extensively for making cutlery and surgical instruments. It is also used in chemical plants because it is extremely durable and is resistant to corrosion even when heated.

Key ideas

1. Iron extracted from the blast furnace is mainly used to produce steel.
2. Steel is an alloy of iron with carbon and/or other metals.
3. Adding different amounts of carbon or metals, and adding different metals to iron produce alloys with different properties.
4. High carbon steels are strong and brittle; low carbon steels are softer and more easily shaped.
5. The uses of steel are given in the table below.

Type of steel	Uses	Special properties
mild steel	car bodies and machinery	hard, strong and malleable
high carbon steel	cutting and boring tools	strong but brittle
stainless steel	equipment in chemical plants, cutlery and surgical instruments	resistant to corrosion

14.5 | Rusting

What is rusting?

It was mentioned in section 14.1 that metals may react with air and water, and corrode. When an object made of iron is exposed to moist air for some time, a reddish-brown substance slowly forms on the surface of the metal. This substance is called **rust** and has the chemical name *hydrated iron(III) oxide*.

Have you ever left safety pins made of iron or steel by the sink and forgotten about them? What do you notice on the pins after a few days?

The process that produces rust is known as **rusting** or the **corrosion of iron**. It is the slow oxidation of iron to form hydrated iron(III) oxide (rust). The equation for rusting is

$$\text{iron} + \text{oxygen} + \text{water} \longrightarrow \underset{\text{(rust)}}{\text{hydrated iron(III) oxide}}$$

$$4Fe(s) + 3O_2(g) + 2xH_2O(l) \longrightarrow 2Fe_2O_3.xH_2O(s)$$

(The x in the chemical formula of hydrated iron(III) oxide indicates the number of water molecules present in the compound.)

What conditions are essential for rusting?

The experiment on page 250 investigates the essential conditions for rusting.

Experiment 2

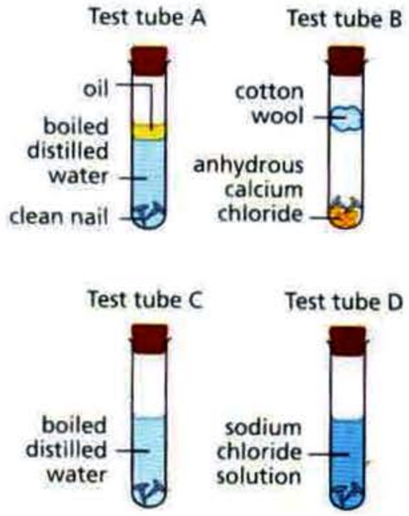


Fig. 14.21 Studying the conditions necessary for rusting

To investigate the conditions necessary for rusting.

Procedure

1. The iron nails are cleaned with sandpaper to remove any rust present.
2. The iron nails are then placed in four different test tubes as shown in Fig. 14.21.
3. After one week, the iron nails are examined and their appearance is noted.

The results of the experiment are summarised in Table 14.9.

Test tube	Conditions		Observation
	Is water present?	Is air present?	
A	yes	no	no rusting
B	no	yes	no rusting
C	yes	yes	nails rusted
D	yes	yes	nails rusted heavily

Table 14.9 Results of experiment 2

ScienceSkills

Scientists have shown that it is oxygen in the air that is essential for rusting. Think of ways to modify experiment 2 to verify this fact.

The results obtained show two important facts about rusting:

1. Both air and water are needed for rusting to occur.
2. The presence of sodium chloride increases the speed of rusting.

Besides sodium chloride, acidic substances such as sulphur dioxide and carbon dioxide also speed up the rusting process. Thus, iron objects near the sea and in industrial areas corrode more rapidly because of the presence of salt and other pollutants in the air.

Rust Prevention

This iron chain rusts rapidly because it is exposed to salt in the air.

Rust is very brittle and flaky. Thus, when iron corrodes, the rusted surface of the metal flakes away. This produces a new metal surface to corrode. Eventually, all of the metal will rust and flake away. There are three general methods of rust prevention — using a protective layer, using a sacrificial metal and using alloys.

How does using a protective layer prevent rusting?

In order to prevent iron from rusting, it has to be kept away from water and oxygen. One general method of 'rust proofing' is to coat the metal object with a protective layer of substance. This includes coating the iron/steel surface with paint or grease, covering it with plastic, electroplating it or dip plating it.

Both electroplating and dip plating involve galvanising, i.e. coating iron/steel with a layer of another metal. Electroplating uses electricity to do this. Dip plating is the process of dipping iron into molten zinc or tin. A thin film of zinc or tin then covers the iron/steel. This layer prevents water and air from coming into contact with the iron or steel surface.

Fig 14.22 Rust prevention using a sacrificial metal

Using a sacrificial metal

Fixing bars of zinc to a ship's hull prevents the ship's steel body from rusting (Fig. 14.22). Heavy blocks of magnesium attached to underground pipes made of iron protects the pipes from rusting (Fig. 14.23). These metals attached are more reactive than iron and will corrode instead of iron. As long as magnesium or zinc is present, iron will not rust. This is called sacrificial protection because the magnesium or zinc is 'sacrificed' to protect the iron or steel.

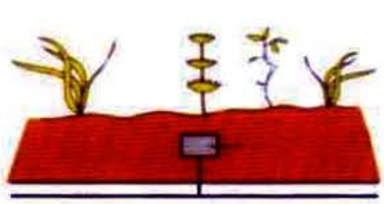


Fig 14.23 Sacrificial protection of underground pipes

Using alloys

The best known rust-resistant alloy of iron is stainless steel. Stainless steel contains nickel and chromium. On exposure to air and moisture, a very hard coating of chromium(III) oxide, Cr_2O_3, forms on the surface of stainless steel, protecting it from further corrosion.

Table 14.10 summarises the common methods of rust prevention.

Method	Where it is used	Comment
painting	large objects like cars, ships and bridges	If the paint on the metal surface is scratched, rusting will take place under the painted surface.
oiling or greasing	tools and machine parts	The protective film of oil or grease gathers dust and must be renewed.
plastic coating	kitchenware such as draining racks	If the plastic layer is torn, the iron starts to rust.
galvanising (zinc-plating)	water buckets, dustbins, 'zinc' roofs, kitchen sinks	The metal does not rust even if the zinc layer is damaged. (This is because zinc is more reactive than iron. So zinc corrodes instead of iron.)
tin-plating	food cans	If the tin layer is scratched, the iron beneath it rusts.
chrome-plating	taps, kettles, bicycle handle bars	This gives a bright shiny finish as well as rust protection.
metal block of zinc or magnesium	underground pipes, ships, legs of steel piers	Magnesium and zinc corrode in place of iron because they are more reactive metals.
stainless steel	cutlery, surgical instruments, pipes in chemical plants	Stainless steel contains chromium and nickel, which do not rust.

Table 14.10 Various methods of rust prevention

Key ideas

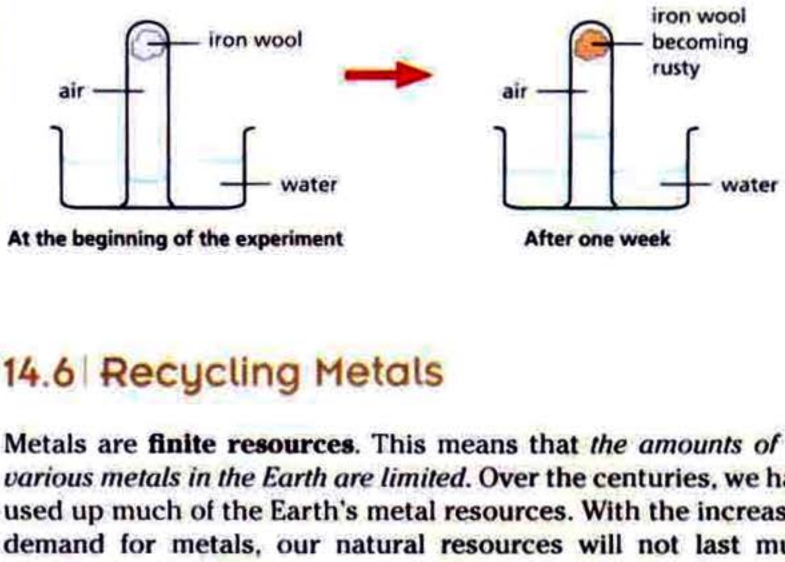
1. Rusting is the oxidation of iron to form hydrated iron(III) oxide. It only occurs in the presence of both oxygen and water.
2. Rusting can be prevented by placing a protective layer over the metal. This includes painting, greasing, applying a plastic coating over the metal surface or galvanising.
3. Placing a protective layer over iron prevents rusting because the protective layer stops oxygen and water from coming into contact with the metal.
4. Rusting can be prevented by sacrificial protection, i.e. attaching a more reactive metal to the iron or steel object. Rusting is prevented because oxygen reacts with the more reactive metal.

Test Yourself 14.5

Questions

1. Describe how rusting is prevented on
   a) the moving parts of a machine.
   b) food cans.
   c) kitchen sink taps.
   d) the parts of a steel pier submerged in water.

2. The diagram shows a simple experiment on rusting. What information can be drawn from the experiment?



14.6 | Recycling Metals

Metals are **finite resources**. This means that *the amounts of the various metals in the Earth are limited*. Over the centuries, we have used up much of the Earth's metal resources. With the increasing demand for metals, our natural resources will not last much longer. Fig. 14.24 shows approximately how long it will take for the reserves of some metals to run out.

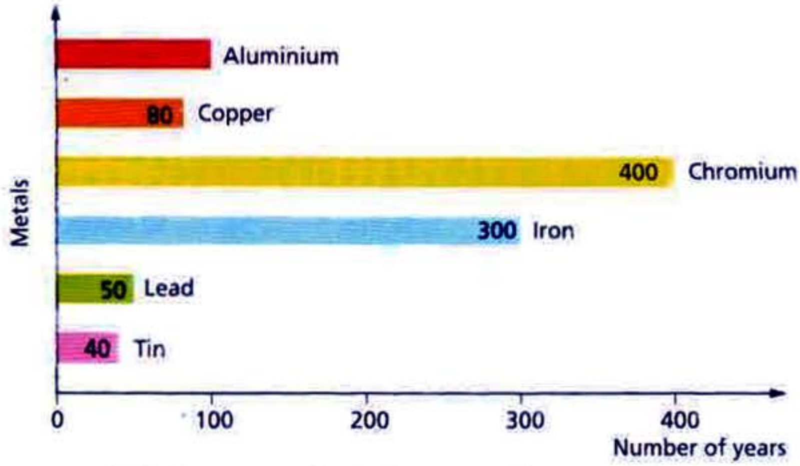


Fig. 14.24 Approximate time for the reserves of some metals to run out

Some metals such as chromium and iron are so abundant that it is difficult to believe they can ever run out. On the other hand, metals such as lead and tin are in very short supply. The reserves of these metals could run out in your lifetime! The world's reserves of raw metals may last longer if

- we find new ore deposits using advanced technology like satellites.
- we find substitutes to replace metals so that we can use the existing metals more sparingly. A good example is the replacement of some metal parts in cars with plastic or ceramic parts. Optical fibres made of special glass are replacing the metal wires used in communications instruments.
- we recycle metals. Metals are readily recycled — old metal objects can be crushed and melted for reuse (Fig. 14.25).

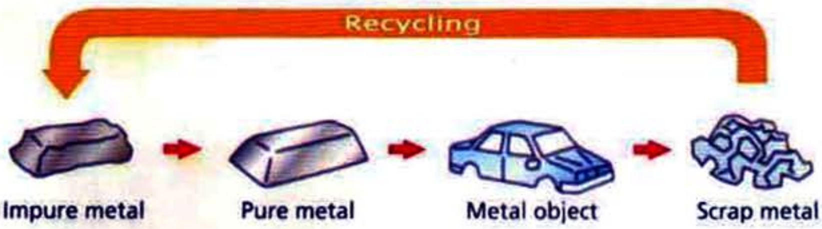


Fig. 14.25 Recycling metals for reuse

Let us now discuss the issues surrounding recycling.

Environmental impact of metal extraction

Before metals can be extracted from their ores, the land is mined for ores. Mined land is usually unsightly and it cannot support plant and animal life.

An enormous amount of waste material is also generated from the extraction process as shown in Fig. 14.26.

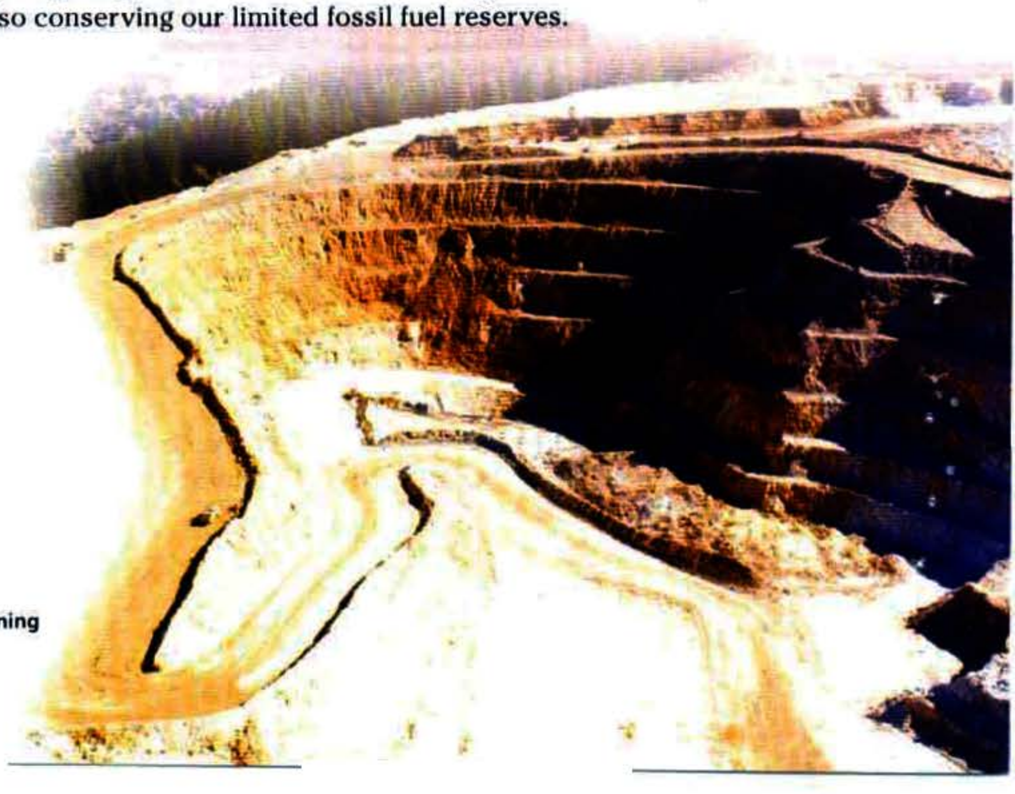
Fig. 14.26 About 90% of a metal ore ends up as waste after the extraction process.

About 90% of the metal ore used for metal extraction turns out to be waste and needs to be disposed of after the metal is extracted. If not disposed of carefully, the waste may leach into soil and nearby water bodies, polluting land and water. To dispose of the waste, huge landfill sites need to be dug.

The smelting of ores also causes more air pollution compared to any other industrial process. In addition, the extraction of metals from their ores requires a continuous supply of energy. The energy is usually generated from burning fossil fuels, which are already in scarce supply.

Advantages of recycling

If metals are recycled and reused, there will be less need to dig for metal ores. With a decrease in mining operations, land will be free for other uses such as agriculture. Air and water pollution will also be greatly reduced. Fewer landfills will be required to dispose of both used metal objects and waste material from metal extraction. This will help save on the cost of building landfill sites. By recycling metals instead of extracting them from ores, we are also conserving our limited fossil fuel reserves.



A type of mining known as open-pit mining leaves very large holes in the ground.

How are some metals commonly recycled?
Lead, iron and aluminium are mainly recovered by scrap metal recycling. Lead is recovered from car batteries. When a car battery no longer works, the lead inside it is reused. A large fraction of iron and steel produced today is also recycled from scrap metal. Aluminium is recycled mainly from drink cans and food containers.

Economic issues of recycling
One problem with recycling metals is that recycling can also be extremely costly. There is the cost of transporting the scrap metal to the processing plant. The different types of metals must be separated before they can be recycled. Additional costs are incurred to sort and clean the scrap metal.

Some metal-producing businesses may decide that the costs are too high and that recycling is not worthwhile. This is especially if the metal to be recycled is not an expensive nor a very valuable one.

Social issues of recycling
We need energy, clean air and water in order to survive. Compared to extracting metals from ores, recycling does not produce as much waste as that may endanger human health. With the increasing human population worldwide, more land will be needed to grow food, rear livestock, and build homes, factories, offices and highways. Building new mines reduces the land available for these other important uses.

Eventually, resources of metal would still be used up. Thus, it would be wise to start developing metal recycling programs and processes that are cost-effective and environmentally friendly now.

Although there are obvious advantages for recycling metals, it will take effort and time for communities and businesses to practise recycling as a way of life. Everyone will need to realise that each of us plays a vital role in conserving our natural resources.

Recycling bins — a common sight along the streets of Singapore

Key ideas

1. Metals are finite resources and need to be conserved.
2. Recycling metals
   - helps us to conserve natural resources, such as land and fossil fuels,
   - helps us to reduce the environmental problems related to extracting metals from metal ores,
   - may save money spent on building landfill sites.

Test Yourself 11.1

Worked Example

Why is it easier to recycle lead than to recycle aluminium?

Answer

Lead is below aluminium in the reactivity series. Lead compounds can be easily reduced to lead using carbon, unlike aluminium. Lead also has a lower melting point than aluminium. Hence, it is easier to separate out lead from its mixtures with other substances for recycling.

(Note: Aluminium forms compounds with strong bonds. Waste aluminium compounds have to be melted in a molten furnace before pure aluminium can be obtained using electrolysis. This is a very costly process.)

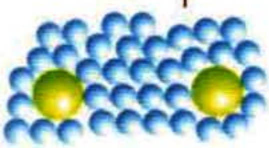
Questions

1. How can other scrap metals be separated from scrap iron?
2. State **three** advantages of recycling aluminium.

Using an industrial-sized magnet to separate scrap metals

Concept Map

Metals — structure — 

- are mixed with other element(s) to form **Alloys** — structure — 
- can be conserved by **Recycling**
- are ordered from most eactive to least reactive in **Reactivity series**
  - potassium
  - sodium
  - calcium
  - magnesium
  - zinc
  - iron
  - lead
  - copper
  - silver
  - gold
- the higher up the metal, the greater its tendency to form → **Positive ions**
- are obtained as follows
  - from ores using electricity
  - from ores by reduction with carbon
  - found uncombined

e.g. to extract iron

Coke (C) + **Limestone** ($CaCO_3$) + **Haematite** (Fe_2O_3 + impurities)

Reactions in blast furnace

- $C(s) + O_2(g) \xrightarrow{heat} CO_2(g)$
- $CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$
- $C(s) + CO_2(g) \xrightarrow{heat} 2CO(g)$
- $Fe_2O_3(s) + 3CO(g) \xrightarrow{heat} 2Fe(l) + 3CO_2(g)$
- $CaO(s) + SiO_2(s) \xrightarrow{heat} CaSiO_3(l)$

→ waste gases + molten iron + slag
(N_2, CO, CO_2)

most metals react

- With water → metal hydroxide + hydrogen
- With steam → metal oxide + hydrogen
- With hydrochloric acid → salt + hydrogen

iron requires both oxygen and water for **Rusting**

prevented by

- Greasing
- Painting
- Plastic coating
- Sacrificial protection

Exercise 14

Foundation

1. Which alloy contains copper as the main constituent?

   A Brass B Pewter
   C Solder D Stainless steel

2. Which metal does not react with dilute hydrochloric acid?

   A Iron B Sodium
   C Zinc D Copper

3. What happens when clean iron filings are placed in copper(II) sulphate solution?

   A Copper(II) ions are oxidised.
   B Iron is oxidised.
   C Iron(III) sulphate is formed.
   D There is no reaction.

4. How is iron extracted from iron(III) oxide in the blast furnace?

   A Oxidise iron(III) oxide with carbon dioxide.
   B Oxidise iron(III) oxide with oxygen.
   C Reduce iron(III) oxide with carbon monoxide.
   D Reduce iron(III) oxide with limestone.

5. Which substance is used to remove impurities from iron ore in the blast furnace?

   A Carbon B Carbon monoxide
   C Limestone D Silica

6. Which pair of metals will slow down rusting when they are in contact with steel?

   A Magnesium and copper.
   B Magnesium and zinc.
   C Zinc and copper.
   D Zinc and lead.

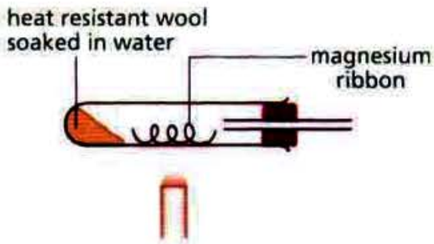
7. a) Place the metals calcium, potassium and zinc in order of chemical reactivity towards water, stating the most reactive metal first.

   b) Write the chemical equation for the reaction between the most reactive metal in (a) and water.

8. a) Steel is an alloy of iron. Explain what you understand by this statement.

   b) Draw diagrams of the structures of iron and steel to show their differences.

9. Magnesium reacts with steam to form magnesium oxide and a gas.



   a) Name the gas formed during the reaction.

   b) Complete the above diagram to show how the gas could be collected.

   c) Describe a test for the gas and state the result you would expect to obtain.

   d) This experiment should not be carried out using potassium. Why?

10. a) Explain what is meant by the term 'recycling'.

    b) Why are copper and lead easy to recycle?

    c) Name the resources that are saved by recycling.

Challenge

1. Metal *X* reacts with water to give a solution *Y*. Solution *Y* gives a white precipitate when carbon dioxide is passed through it. What is metal *X*?

   A Calcium B Lead
   C Nickel D Potassium

2. Which statement about the rusting of iron is not correct?

   A Rusting is a redox process.
   B Rusting is accelerated by the presence of carbon dioxide.
   C Rusting requires both oxygen and water to be present.
   D The composition of rust is $Fe_3O_4.xH_2O$.

3. Four metals, *A*, *B*, *C* and *D*, are tested with water and with dilute hydrochloric acid. The table below shows the results of the experiment.

Metal	Reaction with water	Reaction with steam	Reaction with HC*l* (aq)
A	×	√	√
B	×	×	√
C	√	√	√
D	×	×	×

a) Place the metals *A*, *B*, *C* and *D* in order, stating the most reactive first.

b) Between which two metals should hydrogen be placed in the series in (a)?

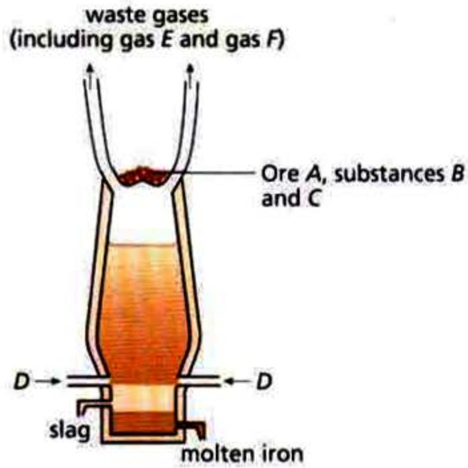
c) Predict the method used for the extraction of metal *C*.

4. a) Nickel is below iron but above lead in the reactivity series. Aqueous solutions of nickel compounds are green. Predict the observations and write the chemical equation for the reaction that takes place when
   i) nickel is added to dilute hydrochloric acid.
   ii) nickel is heated in steam.

b) Predict how you would expect nickel to react when
   i) it is heated with magnesium oxide.
   ii) a piece of nickel is added to copper(II) nitrate solution.

c) i) From the reactions in (b)(i) and (b)(ii), state which of these metals, nickel, magnesium and copper, has the highest tendency to form its positive ions?
   ii) Based on their tendency to form positive ions, deduce the positions of magnesium and copper with respect to nickel in the reactivity series.

5. The diagram below shows the manufacture of iron from ore *A* in a blast furnace.



Substance *B* burns in the furnace to produce gas *E*, which is reduced by more *B* to give gas *F*. This gas reduces the iron ore to iron.

Substance *C* decomposes to form carbon dioxide and a white solid *G*, which reacts with silica to form a molten slag containing calcium silicate. Molten iron and slag settle to the base of the furnace.

a) Identify substances *A* – *G*.

b) Write the equation for the reaction between *G* and silica to form calcium silicate.

c) Select one redox reaction that occurs in the blast furnace. Write the equation for this reaction and identify the reducing agent used.

d) If the haematite is wet, the waste gases also contain hydrogen.
   i) Explain, with an equation, how hydrogen is formed.
   ii) Why is the formation of hydrogen in the furnace dangerous?

e) Iron is often alloyed to reduce rusting. Name one element used for this purpose.

6. You are asked to use simple apparatus to investigate the effect of sulphur dioxide on the corrosion of copper, magnesium and zinc.

a) Why must a control be set up for the experiment?

b) How do you know if corrosion has occurred?

c) Predict the metal that will corrode most easily. Explain your answer.

Chemistry Today

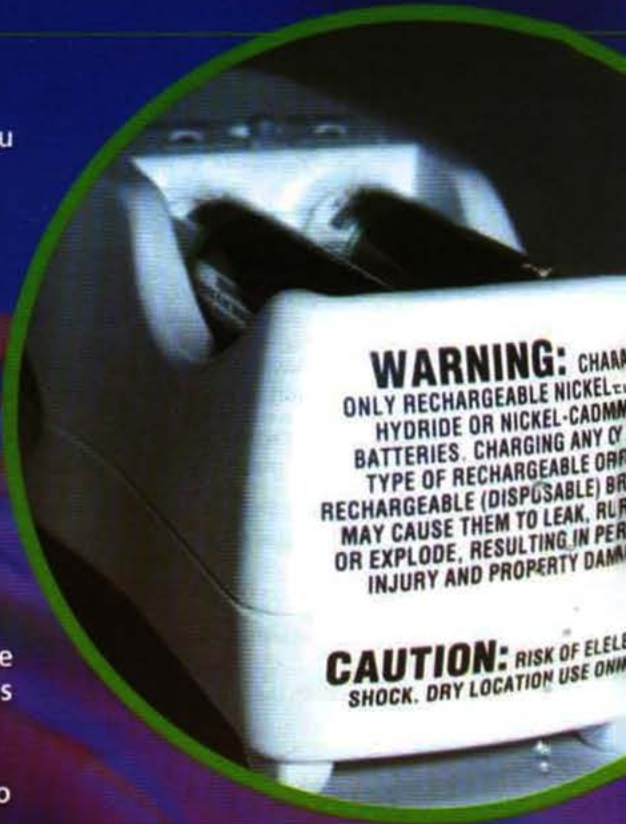
You may have bought rechargeable batteries before. Have you ever read the small print on these batteries? It says 'Please dispose of these wisely due to their cadmium content.' Cadmium, a metal, is a hazardous waste because it is toxic.

The nickel-cadmium battery (commonly abbreviated NiCd or NiCad) is a popular rechargeable battery used in toys, cordless phones and power tools.

Most NiCd batteries are sealed inside these appliances. When the appliances are worn out, they are thrown away. Many people do not realise that by throwing away these appliances, they are contributing to toxic waste. NiCd batteries end up in landfill sites where their casings gradually deteriorate. The cadmium then leaches into the soil and eventually, the water supply. If the batteries happen to be incinerated with rubbish, the ash produced is also a hazardous waste.

Over 1.5 billion NiCd batteries are sold worldwide annually, so the problem caused by disposal is enormous. However, many manufacturers are unwilling to recycle the batteries because the costs are very high and there is very little profit to be made.

NiCd batteries are highly efficient and have a very long shelf life. Thus, scientists are working to invent alternatives that work as well as NiCd batteries but are non-toxic. The nickel metal hydride cell and lithium ion batteries are potential alternatives. Both these types of cells do not contribute to toxic waste. In Europe, the use of cadmium in electrical equipment was banned in December 2004.



Recharging nickel-cadmium batteries

CRITICAL THINKING

You have been asked to reduce the toxicity of batteries. You have thought of four possible solutions:

- Redesigning batteries to reduce the toxic parts
- Replacing batteries with less toxic constituents
- Reducing the number of batteries thrown away by extending battery life
- Recycling

Discuss the arguments for and against each solution.

Chapter 15
Electrolysis

Hydrogen and oxygen gas bubbles being produced when an electric current is passed through water.

In 1800, two Englishmen, William Nicholson and Anthony Carlisle, used an electric current to decompose water into its elements, hydrogen and oxygen. The process of using an electric current to break down water into its components is an example of **electrolysis**.

Electrolysis is an area of **electrochemistry**, which studies chemical reactions that involve electricity. In electrolysis, electrical energy is converted to chemical energy in order for chemical reactions to occur. Another area of electrochemistry has to do with chemical cells or simple cells. In simple cells, chemical energy is converted to electrical energy by means of chemical reactions. In this chapter, you will learn about both electrolysis and simple cells.

Chapter Outline

15.1 | Introducing Electrolysis

Electrolysis is *the process of using electricity to break down or decompose a compound.* The compound is usually molten or dissolved in water. Electrolysis is important for extracting useful pure elements from compounds.

Electrolysis takes place in an **electrolytic cell**, as shown in Fig. 15.1. The electrolytic cell works like an electrical circuit and has these main components — a battery, **electrodes** and an **electrolyte**.

Electrolyte:
- a molten ionic compound or an aqueous solution that conducts electricity.
- dissociates to form positive ions (cations) and negative ions (anions).
- ions present in the electrolyte allow electricity to flow through it.
- examples are dilute sulphuric acid, molten sodium chloride and copper(II) sulphate solution.

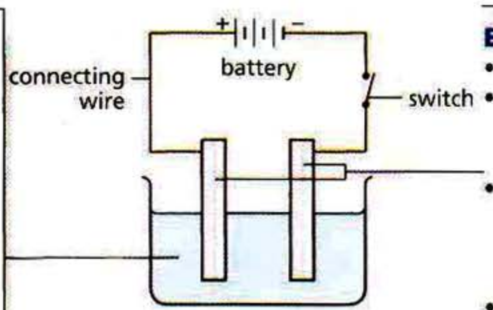


Fig. 15.1 An electrolytic cell

Electrodes:
- conduct electric current.
- usually carbon (i.e. graphite) rods, or metal plates
- electrode that is *connected to the positive terminal of the battery* is called the **anode**.
- electrode that is *connected to the negative terminal of the battery* is called the **cathode**.

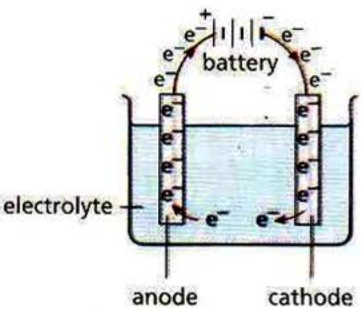


Fig. 15.2 The flow of electrons to and from the battery in an electrolytic cell

How does an electrolytic cell work?
When the circuit is complete in an electrolytic cell, the battery acts as an 'electron pump', drawing electrons away from the anode, which becomes positively charged. These electrons are supplied to the cathode, which becomes negatively charged.

Therefore, electrons enter the positive terminal of the battery and are 'pumped out' at the negative terminal (Fig. 15.2).

What happens when electricity is passed through an electrolyte?
When electricity is passed through an electrolyte, chemical reactions take place at the electrodes, and the electrolyte is decomposed. The reactions taking place at the electrodes are called **electrolytic reactions**.

Chem-Aid

Substances that do not conduct electricity under any condition are called **non-electrolytes**. Examples are sulphur, sugar, pure water and organic compounds such as ethanol.

How can electrolysis account for the existence of mobile ions in molten or aqueous ionic compounds?
Electrolytes only conduct electricity because they contain charged particles (i.e. ions) that are mobile. Since molten or aqueous ionic compounds can act as electrolytes, they must contain mobile ions.

In the solid state, ionic compounds do not conduct electricity and thus cannot act as electrolytes. This is evidence that the positive ions and negative ions in solid ionic compounds are held fixed in a lattice. They are unable to move freely and thus, unable to conduct electricity.

What are the differences between how elements (metals and graphite) and compounds (electrolytes) conduct electricity?
Metals, carbon in the form of graphite and ionic compounds (when molten or aqueous) are all conductors of electricity. Both metals and graphite are known as **electrical conductors** whereas ionic compounds are called **electrolytic conductors**. Table 15.1 shows two major differences between electrical and electrolytic conduction.

	Electrical conduction by metals and graphite	Electrolytic conduction by electrolytes
Method of conduction	Electricity is conducted by the flow of electrons from one end of the conductor to the other end.	Electricity is conducted by the movement of positive ions and negative ions across the (molten and aqueous) electrolyte.
Effect of conduction	Metals and graphite remain unchanged chemically when an electric current flows through them.	The electrolytes are decomposed to form new substances when they conduct electricity.

Table 15.1 The differences between electrical and electrolytic conduction

15.2 | Explaining Electrolysis

The process of electrolysis involves three things: the external circuit, the reactions within the electrolyte, and the reactions at the surface of the electrodes (Fig. 15.3).

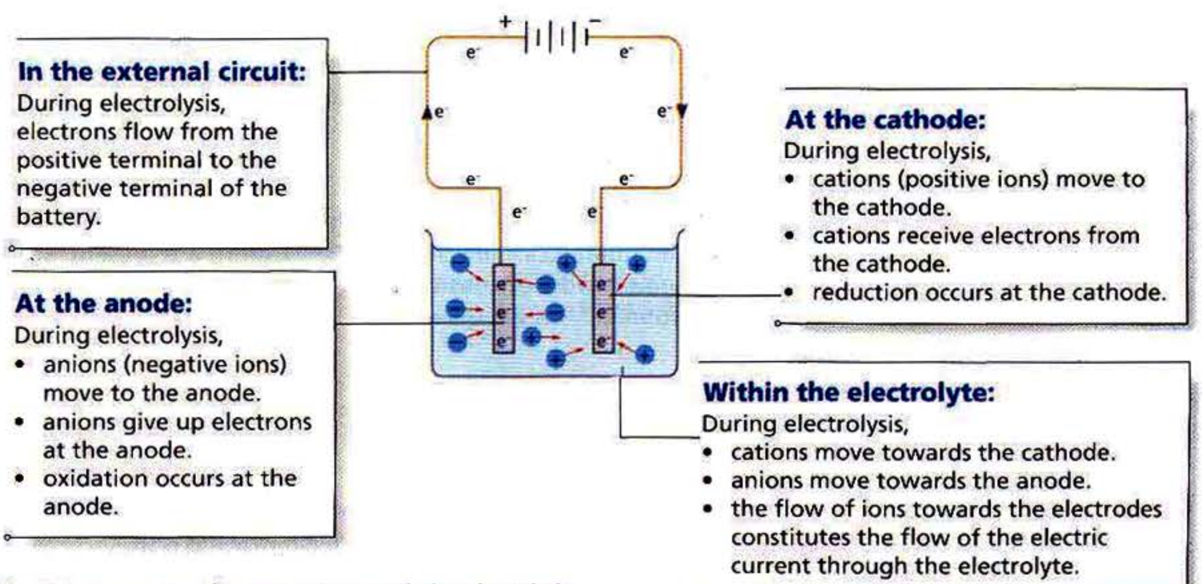
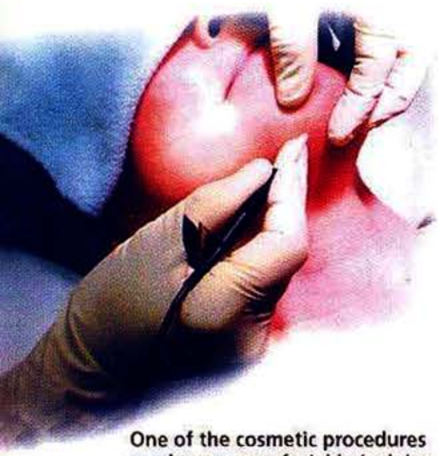


Fig. 15.3 Movement of ions and electrons during electrolysis

Electrical energy is converted to chemical energy as chemical reactions (redox reactions) take place at the electrodes.

What happens to the ions at the electrodes?
As shown in Fig. 15.3, cations (positive ions) receive electrons from the cathode while anions (negative ions) give up electrons to the anode. *The process of gaining or losing electrons at the electrodes* is called **discharge**. When ions are discharged at the electrodes, they form atoms or molecules.



One of the cosmetic procedures used to remove facial hair claims to be electrolysis. Find out how this procedure works and decide if you agree with this claim.

Key ideas

1. The decomposition of a compound by electricity is called electrolysis.
2. Electrolysis takes place in an electrolytic cell. An electrolytic cell consists of a battery, electrodes and an electrolyte.
3. A cathode is a negatively charged electrode. An anode is a positively charged electrode.
4. An electrolyte is a compound that conducts electricity in the molten state or in aqueous solution.
5. An electrolyte conducts electricity because it dissociates into positive ions called cations and negative ions called anions.
6. During electrolysis, cations move towards the cathode while anions move towards the anode.
7. Redox reactions take place at the electrodes during electrolysis. Oxidation occurs at the anode and reduction at the cathode.

Test Yourself 15.1

Worked Example

What is the equation for the reaction that takes place at the negative electrode (cathode) during the electrolysis of molten calcium chloride?

A $Ca^{2+} + 2e^- \longrightarrow Ca$ B $Ca \longrightarrow Ca^{2+} + 2e^-$
C $Ca^{2+} + e^- \longrightarrow Ca^+$ D $2Cl^- \longrightarrow Cl_2 + 2e^-$

Thought Process

Cations move to the cathode where they gain electrons and are reduced. In molten calcium chloride, Ca^{2+} cations are present. Each Ca^{2+} ion gains two electrons to form a calcium atom. This is shown by the equation $Ca^{2+} + 2e^- \longrightarrow Ca$.

Answer

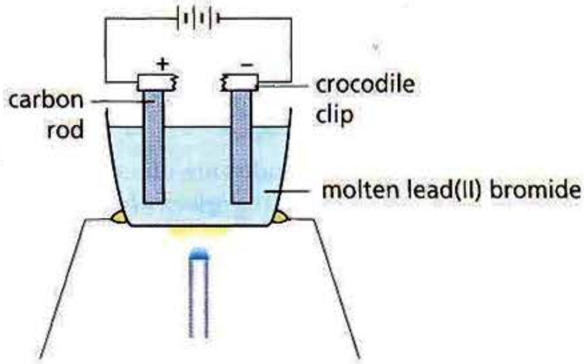
A

Questions

1. Liquid hydrogen chloride is not an electrolyte but an aqueous solution of hydrogen chloride is. Explain why this is so.

2. In the set-up shown, molten lead(II) bromide is electrolysed.

   Copy the figure and
   a) show the direction of the electron flow.
   b) show the movement of lead(II) ions and bromide ions.
   c) identify the electrodes where oxidation and reduction take place.



15.3 | Electrolysis of Molten Ionic Compounds

Some ionic compounds are **binary** compounds. A binary compound is a compound containing only two elements. A binary compound often contains a metal cation and a non-metal anion. When such a molten binary compound undergoes electrolysis, a metal and a non-metal are formed as products.

Sodium chloride is made of only two elements. It is a binary compound.

Electrolysis of Molten Sodium Chloride

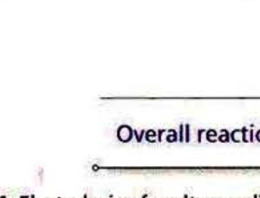
Sodium chloride is an example of a binary compound. It contains sodium ions (Na^+) and chloride ions (Cl^-). When molten sodium chloride is electrolysed (Fig. 15.4), greenish-yellow fumes of chlorine are produced at the anode and greyish molten sodium is produced at the cathode. This method is commonly used to obtain the metal sodium.

What happens when molten sodium chloride is electrolysed?
When solid sodium chloride is heated strongly, it melts at 801 °C. The molten sodium chloride contains mobile Na^+ ions and Cl^- ions. Fig. 15.4 shows the chemical reactions occurring at the electrodes during electrolysis.

At the anode:
Cl^- ions, being negatively charged, are attracted to the anode.

Each Cl^- ion gives up one electron to form a chlorine atom. It is oxidised. Two chlorine atoms then combine to form a chlorine molecule.

$Cl^-(l) \longrightarrow Cl(l) + e^-$
$2Cl(l) \longrightarrow Cl_2(g)$



At the cathode:
Na^+ ions, being positively charged, are attracted to the cathode.

Each Na^+ ion gains one electron to form a sodium atom. It is reduced.

$Na^+(l) + e^- \longrightarrow Na(l)$

Overall reaction: $2NaCl(l) \xrightarrow{\text{electrolysis}} 2Na(l) + Cl_2(g)$

Fig. 15.4 Electrolysis of molten sodium chloride

The lead in pencils contains graphite and can be used as inert electrodes in a simple set-up for electrolysis.

Key ideas

1. The electrolysis of molten sodium chloride produces sodium at the cathode and chlorine at the anode.
2. Inert electrodes do not take part in any chemical reaction during electrolysis.

Why are carbon rods used in the electrolysis of molten sodium chloride?

The chlorine produced during electrolysis is very reactive. **Inert electrodes** such as carbon electrodes are used to prevent reactions from occurring between chlorine and the electrode.

Inert electrodes are electrodes that *do not take part in any chemical reaction during electrolysis*. They simply provide the surface for electron transfer to occur during electrolytic reactions. Inert electrodes are used to prevent reactions from occurring between the electrodes and the products of electrolysis. Carbon and platinum electrodes are considered to be inert electrodes because they are rarely involved in electrolytic reactions.

Test Yourself 15.2

Questions

1. Write an equation, including state symbols, to represent the overall reaction that occurs when molten copper(II) bromide is electrolysed.
2. Molten lithium hydride, LiH, is electrolysed using carbon electrodes.
   a) Name the ions present in molten lithium hydride.
   b) Predict the products obtained at the anode and cathode during electrolysis.
   c) Write the equations for the reactions occurring at the carbon electrodes.

15.4 | Electrolysis of Aqueous Solutions of Compounds

As mentioned earlier, electrolysis can be used to decompose molten compounds or compounds in aqueous solutions. The reactions that occur at the electrodes are rather easy to work out for the electrolysis of molten compounds. It is slightly more complex to predict the reactions for the electrolysis of aqueous solutions. In this section, you will learn why this is so and how to work out the reactions correctly.

An aqueous solution of a compound is really a mixture of two electrolytes. For example, an aqueous solution of copper(II) sulphate contains two electrolytes, namely copper(II) sulphate and water. The aqueous solution, therefore, contains copper(II) ions (Cu^{2+}) and sulphate ions (SO_4^{2-}), and also small amounts of hydrogen ions (H^+) and hydroxide ions (OH^-) from the dissociation of water molecules.

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

These ions compete with the ions from copper(II) sulphate for discharge at the electrodes.

In general, when an aqueous solution of an ionic compound is electrolysed, a metal or hydrogen is produced at the cathode. At the anode, a non-metal, for example oxygen or a halogen, is given off. Let us see if this is the case in the electrolysis of dilute sodium chloride solution.

Electrolysis of Dilute Sodium Chloride Solution

An aqueous solution of sodium chloride contains four different types of ions. They are
- ions from sodium chloride — $Na^+(aq)$ and $Cl^-(aq)$
- ions from water — $H^+(aq)$ and $OH^-(aq)$

What ions are present in an aqueous solution of sodium chloride?

When dilute sodium chloride solution is electrolysed using inert electrodes, the Na^+ and H^+ ions are attracted to the cathode. The Cl^- and OH^- ions are attracted to the anode. Fig. 15.5 shows the reactions that occur and the products obtained at the electrodes.

At the cathode:
The H^+ and Na^+ ions are attracted to the platinum cathode. H^+ ions gain electrons from the cathode to form hydrogen gas.

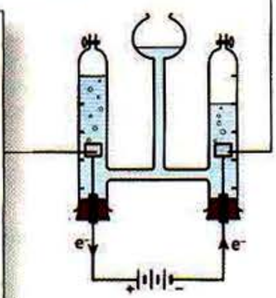
$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

Na^+ ions remain in solution.

At the anode:
OH^- and Cl^- ions are attracted to the platinum anode. OH^- ions give up electrons to the anode to form water and oxygen gas.

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

Cl^- ions remain in solution.



Summary:
The overall reaction is

$$2H_2O(l) \xrightarrow{\text{electrolysis}} 2H_2(g) + O_2(g)$$

Since water is being removed (by decomposition into hydrogen and oxygen), the concentration of sodium chloride solution increases gradually.

Fig. 15.5 Electrolysis of dilute sodium chloride solution

The overall reaction shows that the electrolysis of dilute sodium chloride solution is equivalent to the electrolysis of water.

Compare the electrolysis of dilute sodium chloride solution and that of molten sodium chloride. Fig. 15.6 shows the ions that are discharged during the electrolysis.

Electrolysis

Molten sodium chloride
Cathode: Na^+ ions discharged
Anode: Cl^- discharged

Dilute sodium chloride solution
Cathode: H^+ ions discharged
Anode: OH^- discharged

Fig. 15.6 Comparing the products of the electrolysis of molten sodium chloride and dilute sodium chloride solution

Na^+ and Cl^- ions are not always discharged even though both electrolytes contain these ions. For example, in the electrolysis of dilute sodium chloride solution, H^+ ions are discharged in preference to Na^+ ions. OH^- ions are discharged in preference to Cl^- ions. Why is one type of cation (or anion) in the electrolyte more readily discharged than another type? We shall study the factors involved.

Chem-Aid

The reactivity series is also known as the electrochemical series.

Reactivity Series and Selective Discharge of Ions

In electrolysis, when more than one type of cation or anion is present in a solution, only one cation and one anion are preferentially discharged. This is known as the **selective discharge** of ions.

How do you predict which ions are discharged in the electrolysis of a compound in aqueous solution?

If inert electrodes are used during electrolysis, the ions discharged and hence the products formed depend on three factors:

1. The position of the metal (producing the cation) in the reactivity series.
2. The relative ease of discharge of an anion.
3. The concentration of the anion in the electrolyte.

The ease of discharge of cations and anions is shown below.

Cations

potassium ion, K^+
sodium ion, Na^+
calcium ion, Ca^{2+}
magnesium ion, Mg^{2+}
zinc ion, Zn^{2+}
iron ion, Fe^{2+}
lead ion, Pb^{2+}
hydrogen ion, H^+
copper ion, Cu^{2+}
silver ion, Ag^+

Anions

chloride ion, Cl^-
bromide ion, Br^-
iodide ion, I^-
hydroxide ion, OH^-

Note: sulphate ions (SO_4^{2-}) and nitrate ions (NO_3^-) will not be discharged during electrolysis.

Fig. 15.7 Relative ease of discharge of ions during electrolysis

Quick Check

Two ionic compounds, zinc sulphate and copper(II) sulphate, are dissolved in water and electrolysed using platinum electrodes.

a) Name the ions present in the solution.
b) Predict which ion will be discharged at the platinum cathode.

Selective discharge of cations during electrolysis

The cations of an element lower in the reactivity series are discharged at the cathode in preference to other cations in the solution. This is because cations of a less reactive element accept electrons more readily. For example, if a solution containing Na^+ and H^+ ions is electrolysed, H^+ ions are discharged in preference to Na^+ ions.

Selective discharge of anions during electrolysis

Sulphate (SO_4^{2-}) and nitrate (NO_3^-) ions remain in the solution and are not discharged during electrolysis. If a solution containing SO_4^{2-}, NO_3^- and hydroxide (OH^-) ions is electrolysed, the OH^- ions will be discharged in preference to SO_4^{2-} and NO_3^- ions. The OH^- ions give up electrons most readily during electrolysis to form water and oxygen.

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

Effect of concentration on selective discharge of anions
An increase in the concentration of an anion tends to promote its discharge. For example, in the electrolysis of concentrated sodium chloride solution, two types of ions are attracted to the anode: Cl^- and OH^- ions. According to their relative ease of discharge, OH^- ions should be discharged preferentially. However, in concentrated sodium chloride solution, Cl^- ions are far more numerous than OH^- ions and so are discharged at the anode instead (see next page).

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$$

What are the general rules for predicting selective discharge?
The following rules can be applied when predicting the products of electrolysis of any aqueous solution (using inert electrodes):

Rule 1	Identify the cations and anions in the electrolyte. Remember that an aqueous solution also contains H^+ and OH^- ions from the dissociation of water molecules.
Rule 2	At the anode, the product of electrolysis is always oxygen unless the electrolyte contains a high concentration of the anions, Cl^-, Br^- or I^- ions.
Rule 3	At the cathode, reactive metals such as sodium and potassium are never produced during electrolysis of the aqueous solution. If the cations come from a metal above hydrogen in the reactivity series, then hydrogen will be liberated. If the cations come from a metal below hydrogen, then the metal itself will be deposited.
Rule 4	Identify the cations and anions that remain in the solution after electrolysis. They form the product remaining in solution. Summarise the reactions. For example, in the electrolysis of dilute sodium chloride solution, Na^+ and Cl^- ions remain in solution after H^+ and OH^- ions have been discharged. Hence, the solution of sodium chloride becomes more concentrated after electrolysis.

Table 15.2 Rules for predicting the products of electrolysis of any aqueous solution (when using inert electrodes)

We shall apply these rules for the electrolysis of aqueous copper(II) sulphate and concentrated sodium chloride solution.

Electrolysis of Aqueous Copper(II) Sulphate Using Inert Electrodes

Copper(II) sulphate solution can be electrolysed using inert platinum electrodes.

During electrolysis, the cathode is coated with a layer of reddish-brown solid copper. The blue colour of the solution fades gradually as more copper is deposited. The resulting electrolyte also becomes increasingly acidic.

Quick Check

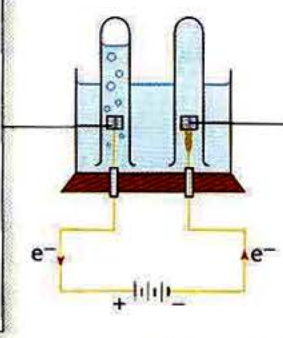
What would you observe when a concentrated solution of hydrochloric acid is electrolysed using platinum electrodes? Explain your observations and write down the equations involved.

Rule 1:
An aqueous solution of copper(II) sulphate contains four types of ions:
Ions from copper(II) sulphate — Cu^{2+} and SO_4^{2-}
Ions from water — H^+ and OH^-

Rule 2: At the anode
OH^- ions and SO_4^{2-} ions are attracted to the anode. OH^- ions give up electrons more readily than SO_4^{2-} ions. Consequently, OH^- ions are preferentially discharged to give oxygen gas.

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

The SO_4^{2-} ions remain in solution.



Rule 3: At the cathode
H^+ ions and Cu^{2+} ions are attracted to the cathode. Copper is lower than hydrogen in the reactivity series. Cu^{2+} ions accept electrons more readily than H^+ ions. As a result, Cu^{2+} ions are preferentially discharged as copper metal (atoms).

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

The H^+ ions remain in solution.

Rule 4: Summary
When aqueous copper(II) sulphate is electrolysed using platinum electrodes, copper metal is deposited at the cathode and oxygen gas is given off at the anode. The overall reaction is

$$2CuSO_4(aq) + 2H_2O(l) \xrightarrow{\text{electrolysis}} 2Cu(s) + O_2(g) + 2H_2SO_4(aq)$$

Fig. 15.8 Electrolysis of aqueous copper(II) sulphate using inert platinum electrodes

Electrolysis of Concentrated Sodium Chloride Solution

The electrolysis of concentrated sodium chloride solution can be carried out in the set-up shown in Fig. 15.9.

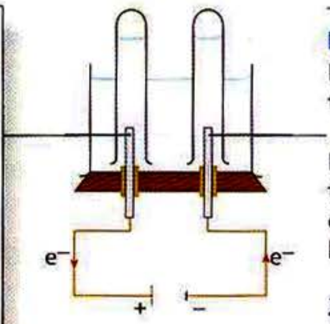
Rule 1: The following ions are present in concentrated sodium chloride solution:
Ions from sodium chloride — Na^+ and Cl^-
Ions from water — H^+ and OH^-

Rule 2: At the anode:
OH^- ions and Cl^- ions are attracted to the anode.

Cl^- ions are more numerous than OH^- ions (see page 269). Consequently, Cl^- ions are discharged as chlorine gas, which bubbles off.

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$$

The OH^- ions remain in solution.



Rule 3: At the cathode:
H^+ ions and Na^+ ions are attracted to the cathode.

H^+ ions accept electrons more readily than Na^+ ions. As a result, H^+ ions are discharged as hydrogen gas, which bubbles off.

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

The Na^+ ions remain in solution.

Rule 4: Summary
One volume of hydrogen gas is given off at the cathode and one volume of chlorine gas is produced at the anode. The resulting solution becomes alkaline because there are more OH^- ions than H^+ ions left in the solution.

Fig. 15.9 Electrolysis of concentrated sodium chloride solution

Electrolysis of Water

Pure water is a poor conductor of electricity because it consists almost entirely of molecules and has very few ions in it. However, if a small amount of an ionic compound or dilute sulphuric acid is added to water, the solution now becomes a good conductor of electricity. The products of the electrolysis of water are always two volumes of hydrogen at the cathode and one volume of oxygen at the anode.

Link

Why is the volume of hydrogen formed always twice that of oxygen during the electrolysis of water? Recall information on the composition of compounds from chapter 4 to explain this.

Key ideas

1. The table below shows the products of electrolysis of various aqueous solutions using inert electrodes.

Aqueous solution	At the cathode	At the anode
NaC*l*(aq), concentrated	hydrogen (1 volume)	chlorine (1 volume)
NaC*l*(aq), dilute (equivalent to the electrolysis of water)	hydrogen (2 volumes)	oxygen (1 volume)
$CuSO_4$(aq)	copper	oxygen

2. When predicting the products of electrolysis of an aqueous solution using inert electrodes:

At the cathode	At the anode
OH^- ions will be discharged as oxygen gas. However, if a solution is concentrated in halide ions, the halogen gas will be produced.	H^+ ions will be discharged producing hydrogen unless the other cation is below hydrogen on the reactivity series.

Try it Out

Electrolysis is not a cost-effective way of producing hydrogen on a large scale, mainly because the electricity needed to run the process is expensive. Think of an alternative, renewable source of energy that is abundant in Singapore that could be used for electrolysis. With the help of information from the Internet, devise a method to electrolyse seawater using this source of energy.

The electrolysis of seawater is equivalent to the electrolysis of water. Hydrogen and oxygen gases are produced.

Test Yourself 15.3

Worked Example

Predict the products obtained at the anode and the cathode when a concentrated solution of copper(II) chloride is electrolysed using carbon electrodes.

Thought Process

Cations present in concentrated copper(II) chloride solution: Cu^{2+}, H^+
Anions present in concentrated copper(II) chloride solution: Cl^-, OH^-
Cu^{2+} will be selectively discharged at the cathode since copper is lower on the reactivity series than hydrogen.
At the cathode: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

According to the relative ease of discharge of anions, OH^- ions would be discharged preferentially over Cl^- ions at the anode. However, in concentrated copper(II) chloride solution, Cl^- ions are far more numerous than OH^- ions. Thus, Cl^- ions are discharged instead.
At the anode: $2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$

Answer

A reddish-brown deposit of copper metal is deposited on the cathode. Greenish-yellow fumes of chlorine gas are produced at the anode.

Questions

1. Predict the substances produced at the cathode and anode when sodium sulphate solution is electrolysed using carbon electrodes. Write an equation for the overall reaction that occurs.

2. What would you observe when a concentrated solution of potassium iodide is electrolysed using platinum electrodes? Briefly explain your answer. Write ionic equations for the reactions that occur at the electrodes.

15.5 | Industrial Applications of Electrolysis

Crude copper extracted from its ore is called 'blister' copper and is about 98% pure copper. Before it can be used in electrical wiring, it must be purified further. Very pure copper (99.9% purity) is obtained from crude copper by electrolysis.

There are many industrial applications of electrolysis. In this section, we shall discuss how electrolysis is used to purify metals (electrolytic purification) and in electroplating.

Electrolytic Purification

In this age of information technology, there is a high demand for very pure metals in printed circuits and for other specialised uses. Electrolysis is used to purify metals, such as in the electrolytic purification of copper.

Earlier, you learnt about the electrolysis of aqueous copper(II) sulphate using inert electrodes. The electrolytic purification of copper also involves the electrolysis of aqueous copper(II) sulphate, but using copper electrodes.

Electrolysis of aqueous copper(II) sulphate using copper electrodes

When copper(II) sulphate is electrolysed using inert electrodes, copper is formed at the cathode and oxygen at the anode. However, if copper electrodes are used, the electrodes may undergo oxidation in preference to other possible reactions. In this case, the electrodes are called **reactive electrodes.**

What happens when aqueous copper(II) sulphate is electrolysed using reactive electrodes?

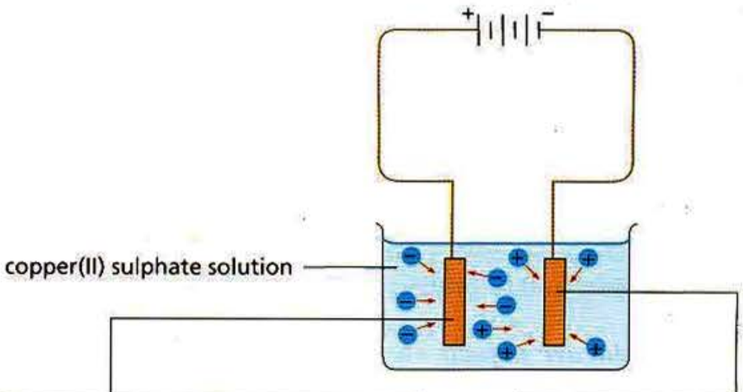
When aqueous copper(II) sulphate is electrolysed using copper electrodes (reactive electrodes),

a) no gas is evolved at the anode,
b) the copper anode slowly becomes smaller,
c) copper is deposited on the cathode, and
d) the blue colour of the aqueous solution remains unchanged.

How do you account for these observations?

Aqueous copper(II) sulphate contains four types of ions:

- Ions from copper(II) sulphate — Cu^{2+} and SO_4^{2-}
- Ions from water — H^+ and OH^-



At the copper anode:

SO_4^{2-} and OH^- ions migrate to the anode. Since copper is not an inert electrode, the following reaction occurs at the anode:

$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

The anode dissolves to form Cu^{2+} ions in aqueous solution.

At the copper cathode:

Cu^{2+} and H^+ ions migrate to the cathode. Here, copper is formed rather than hydrogen gas because Cu^{2+} ions accept electrons more readily than H^+ ions do.

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

The cathode becomes coated with a layer of reddish-brown copper.

Fig. 15.10 Electrolysis of aqueous copper(II) sulphate using copper electrodes

What is the net result of electrolysis of aqueous copper(II) sulphate using copper electrodes?
The net result is that copper is transferred from the anode to the cathode. The copper cathode slowly increases in mass while the anode decreases in mass. The colour and concentration of the copper(II) sulphate solution remain unchanged. This is because the Cu^{2+} ions that get discharged come mainly from the copper anode. There is no effective loss of Cu^{2+} ions from the copper(II) sulphate solution.

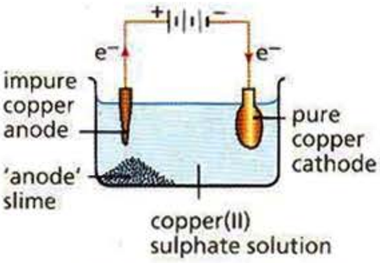


Fig. 15.11 Purification of copper

Purification of copper
To refine copper, impure copper is used as the anode of the electrolytic cell. The cathode is a thin sheet of pure copper and the electrolyte is aqueous copper(II) sulphate. During electrolysis, the impure copper anode dissolves and impurities such as silver and platinum fall to the bottom of the cell. In a finely divided state, they form an 'anode slime'. A layer of pure copper is deposited on the cathode.

Electroplating

The process of depositing a layer of metal on another substance using electrolysis is called **electroplating.**


Food cans are tin-plated by electrolysis. Tin-plated food cans rust less easily.

What are the uses of electroplating?
Electroplating is used to coat a metal object with another metal to get a good decorative finish and to prevent rusting. Food cans are electroplated with tin for these reasons. Silver-plating or gold plating is used to coat a relatively cheap metal to make it look more expensive. Chrome-plating is also used to beautify a metal object and to protect it from corroding.

How are objects electroplated?
Electroplating, like electrolytic refining, also makes use of electrolysis.

The object to be plated is made the cathode of an electrolytic cell and the anode is the source of the plating metal. The electrolyte is an aqueous solution of a salt of the plating metal. The net result is the transfer of the plating metal from the anode to the cathode.

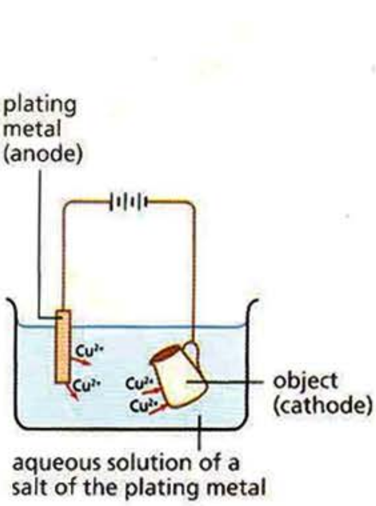


Fig. 15.12 Electroplating

Copper plating
Apparatus similar to that shown in Fig. 15.12 is used to coat a metal object with a thin layer of copper. The anode is pure copper, often called the plating copper. The metal object to be copper-plated is made the cathode and the electrolyte is aqueous copper(II) sulphate.

At the anode: $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

At the cathode: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

Key ideas

1. The table below compares the products obtained from the electrolysis of aqueous copper(II) sulphate using inert (platinum) and reactive (copper) electrodes.

Electrolyte	Cathode	Anode	Reaction at cathode	Reaction at anode
aqueous copper(II) sulphate	platinum	platinum	copper deposited	oxygen liberated
	copper	copper	copper deposited	copper anode dissolves

2. The electrolysis of aqueous copper(II) sulphate using copper electrodes is used for refining impure copper and copper plating objects.
3. Electroplating of metals helps to prevent their corrosion and improve their appearance.

TidBit

What conditions favour good quality copper-plating? There are three conditions that must be satisfied for a thin layer of copper to adhere well to the surface being coated:

1. The metal object to be plated must be clean and free of grease.
2. The concentration of Cu^{2+} ions in the electrolyte must be low.
3. The electric current must not be too large, otherwise the coating layer will form too rapidly and peel off easily.

A copper-plated hook

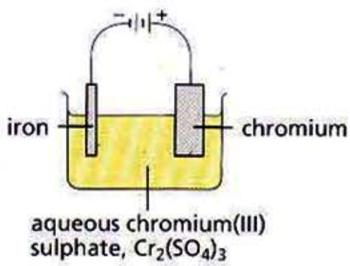
ScienceSkills

You can try copper-plating small iron or steel objects in the school laboratory. Predict which factors could affect the mass of copper deposited on the object you are plating. Explain your answer briefly. What tests would you carry out to confirm your predictions?

Test Yourself 15.4

Worked Example

The electrolytic cell shown below is a simple representation of an industrial process.



a) Identify the anode and cathode.
b) Describe the net result of the electrolysis.
c) Suggest which industrial process this cell represents.

Thought Process

a) The anode is the electrode connected to the positive terminal of the battery; the cathode is the electrode connected to the negative terminal of the battery.
b) Cr atoms from the anode gets oxidised to Cr^{3+} ions. These cations move through the electrolyte to the cathode. At the cathode, Cr^{3+} ions get reduced to Cr atoms, effectively coating iron with a layer of chromium.
c) Iron is plated with chromium.

Answer

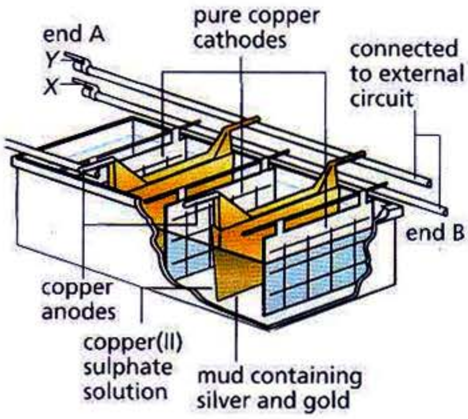
a) Anode: chromium electrode.
   Cathode: iron electrode.

b) The net result is the transfer of chromium from the anode to the iron cathode.

c) Chromium-plating of iron.

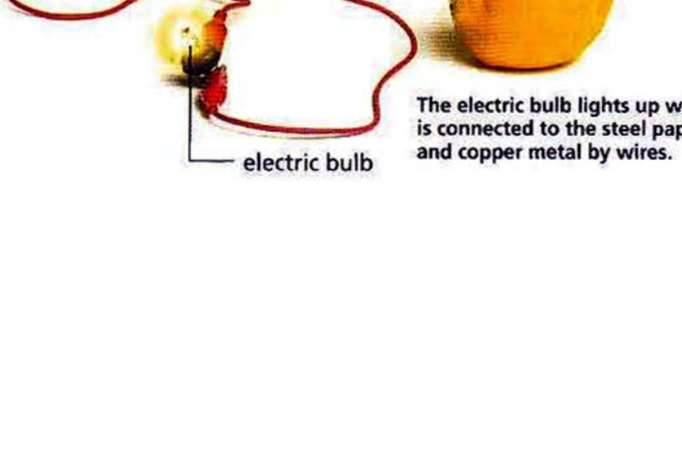
Questions

1. The diagram below shows an industrial process.
   a) Examine the diagram carefully. Is the anode *X* or *Y*? Explain your answer.
   b) What is the industrial application of this process?
   c) In which direction do electrons flow in rod *X* and rod *Y*, towards end A or end B?



2. a) Draw a labelled diagram to show how a steel spoon can be electroplated with chromium.
   b) Zinc is cheaper than tin but zinc metal is never used to electroplate food cans. Why is this so?

15.6 | Simple Cells



The electric bulb lights up when it is connected to the steel paper clip and copper metal by wires.

In electrolysis, batteries supply electrical energy for chemical reactions to occur. This way, electrical energy is converted to chemical energy. Study the picture. The lemon with the two metals stuck into it acts like a battery. It supplies electrical energy to light up the bulb.

For a battery to produce electrical energy, a chemical change must first take place inside the battery. The chemical change produces chemical energy that is converted to electrical energy. A battery is made up of **simple cells**. For example, a 12 V car battery is made up of six simple cells joined together in series. It is in these simple cells that chemical reactions occur to eventually produce electricity.

What is a simple cell?

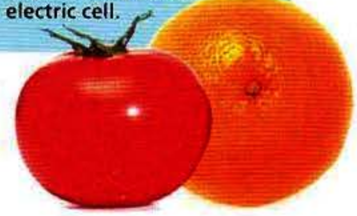
A simple cell is a *device which converts chemical energy to electrical energy*. It is also known as an **electric cell**. It is made by placing two different metals in contact with an electrolyte. The metals act as the electrodes for the simple cell.

For example, when a zinc plate and a copper plate are placed in dilute sulphuric acid and connected by wires, a chemical reaction occurs to produce an electric current that flows through the wires. A potential difference is set up between the metal plates. This potential difference or voltage can be registered by a voltmeter (Fig. 15.13).

Try it Out

A simple electric cell can be made by inserting two electrodes made from different metals into a fruit (e.g. a grapefruit or a tomato), as shown on page 276. Try and make such a cell, and see if your cell car be used to run an electric clock or a small motor.

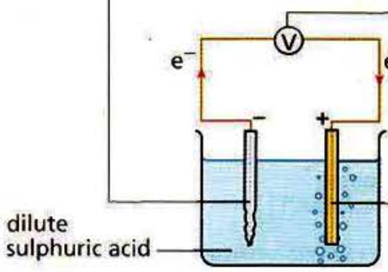
Find out
a) which fruit and metal electrodes will produce the strongest current.
b) if vegetables can be used to produce a simple electric cell.



How is the electric current produced?

At the zinc electrode:
Oxidation occurs. Zinc atoms give up electrons and go into solution as zinc ions.

$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$ (oxidation)



The electrons leave the zinc plate, pass through the voltmeter and enter the copper plate.

At the copper electrode:
The electrons are taken up by the positively charged hydrogen ions to form hydrogen gas. This is a reduction reaction.

$2H^+(aq) + 2e^- \longrightarrow H_2(g)$ (reduction)

The electrode from which electrons flow out of is the negative electrode. The electrode into which the electrons flow is the positive electrode. In this simple cell, zinc acts as the negative electrode and copper the positive electrode. The electrolyte is dilute sulphuric acid. The electrons which flow in the external circuit constitute the electric current.

Fig. 15.13 Production of electric current in a simple cell

In the zinc-copper cell, the zinc plate dissolves in the acid and bubbles of hydrogen gas form at the copper electrode. The overall cell reaction for this simple cell is:

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Oxidation and reduction occur together in order to cause the flow of electrons in a simple cell. Thus, electrical energy is produced by redox reactions in a simple cell.

How do you determine if a particular electrode is positive or negative in a simple cell?

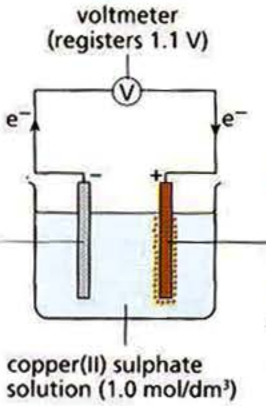
Recall from chapter 14 that more reactive metals tend to give up electrons and form ions more readily than less reactive metals. Thus, in a simple cell, the flow of electrons is always from the more reactive metal to the less reactive metal. The more reactive metal becomes the negative electrode and the less reactive metal the positive electrode.

What happens in a zinc-copper cell when the electrolyte is copper(II) sulphate?

Fig.15.14 shows a zinc-copper cell that uses a 1.0 mol/dm^3 copper(II) sulphate solution which acts as the electrolyte.

At the zinc electrode:
Zinc dissolves to form zinc ions.

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$



At the copper electrode:
Copper(II) ions from the solution receive electrons to form copper.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

Fig. 15.14 A zinc-copper simple cell with aqueous copper(II) sulphate as the electrolyte

The ionic equation for the overall reaction is obtained by adding the half-equations for the reactions at the electrodes.

Overall ionic equation: $Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

Simple Cells and Reactivity Series

You have learnt about the reactivity series in chapter 14, where the reactions of metals with water and acids are used to place the metals in order of their reactivity.

Simple cells can also be used to determine the relative positions of metals in the reactivity series. This is because the amount of electrical energy produced in a simple cell is determined by how far apart the metals used (as electrodes) are in the reactivity series.

By using different metals in the (simple cell) set-up shown in Fig.15.15, different voltages are produced (Table 15.3).

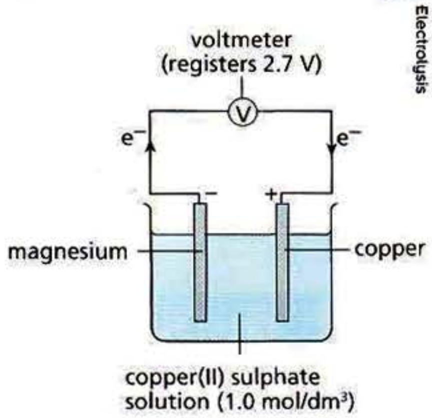


Fig. 15.15 Determining the voltage of a simple cell

Metal electrodes	Voltage (V)
magnesium/copper	2.7
zinc/copper	1.1
iron/copper	0.8
lead/copper	0.5
copper/copper	0.0

Table 15.3 Voltages produced by a simple cell using different metals as electrodes

The voltage of a magnesium-copper cell is 2.7 V. If magnesium is replaced by zinc, a less reactive metal, the voltage decreases to 1.1 V. This shows that the further apart the two metals are in the reactivity series, the greater the voltage produced. No current will flow if both electrodes are made of the same metal.

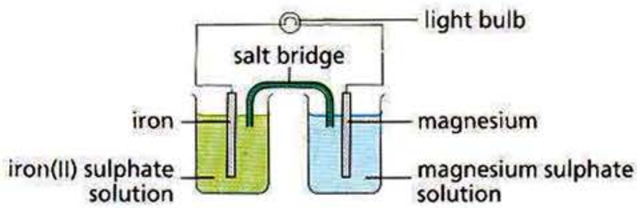
Key ideas

1. A simple cell is made by placing two different metals in an electrolyte.
2. In a simple cell, the electrons flow from the more reactive metal to the less reactive metal.
3. The further apart the two metals are in the reactivity series, the greater the cell voltage produced.

Test Yourself 15.5

Questions

1. Consider the diagram below. Describe the direction of electron flow.



2. Each of the following simple cells has a light bulb in the circuit.

Cell	Electrodes used	Electrolyte used
A	Cu/Mg	distilled water
B	Zn/Fe	distilled water
C	Cu/Mg	dilute sulphuric acid
D	Zn/Fe	dilute sulphuric acid

a) Which cell will cause the bulb to shine most brightly?
b) Briefly explain your reasoning.

TidBit

The electric cell on the left is made up of two **half-cells**, joined by a **salt bridge**. Each half-cell consists of an electrode surrounded by an electrolyte.

A salt bridge completes the circuit, while keeping the electrolytes separate. The salt bridge can be either a glass tube filled with gel containing an inert electrolyte, or a piece of filter paper soaked in the inert electrolyte, usually potassium nitrate.

Concept Map

Simple cell → Chemical energy → Electrical energy

Electrolysis → Electrical energy → Chemical energy

Simple cell
- Electrons flow from the more reactive metal to the less reactive metal.
- The further apart the two metals used as electrodes are in the reactivity series, the higher the voltage produced.

Electrolysis: Passing electricity through an electrolyte (molten/aqueous ionic compound) to decompose it

Ease of discharge of ions during electrolysis

Ease of discharge increases (downwards)

Cations	Anions
K^+	Cl^-
Na^+	Br^-
Ca^{2+}	I^-
Mg^{2+}	OH^-
Zn^{2+}	
Fe^{2+}	
Pb^{2+}	
H^+	
Cu^{2+}	
Ag^+	

(*Note: SO_4^{2-} and NO_3^- are not discharged.)

factors affecting selective discharge:
- position of metal in the reactivity series
- ease of discharge of anion
- concentration of anion in the electrolyte

for predicting

Products of electrolysis

	Electrolyte	Electrodes	at anode (+)	at cathode (–)
Using inert electrodes	molten $NaCl(l)$	carbon	$Cl_2(g)$	$Na(l)$
Using inert electrodes	concentrated $NaCl(aq)$	carbon	$Cl_2(g)$	$H_2(g)$
Using inert electrodes	dilute $NaCl(aq)$	carbon	$O_2(g)$	$H_2(g)$
Using inert electrodes	$CuSO_4(aq)$	platinum	$O_2(g)$	$Cu(s)$
Using reactive electrodes	$CuSO_4(aq)$	copper	Cu anode dissolves	$Cu(s)$
Using reactive electrodes	$CuSO_4(aq)$	Cu (anode) C (cathode)	Cu anode dissolves	$Cu(s)$

industrial applications

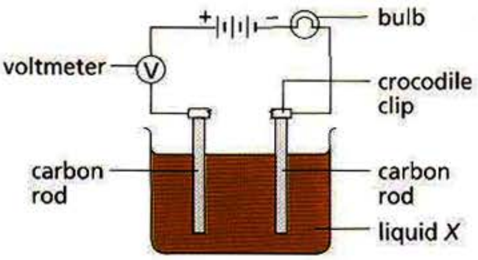
Copper plating
Anode: pure copper
Cathode: object to be plated

Purification of copper
Anode: impure copper
Cathode: pure copper

Exercise 15

Foundation

1. When an experiment was set up as shown below, the voltmeter showed a reading although there were no reactions occurring at the electrodes.



What is liquid *X*?

A Mercury
B Paraffin
C Molten sodium bromide
D Dilute aqueous sodium bromide

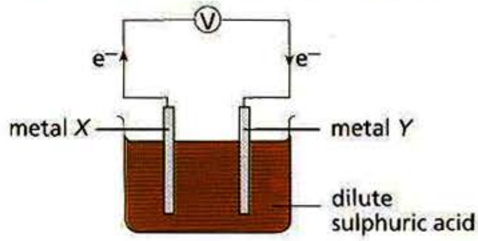
2. Aqueous copper(II) sulphate is electrolysed using carbon electrodes. What would you expect to see during electrolysis?

	At the cathode (–)	At the anode (+)
A	pink solid forms	anode dissolves
B	colourless gas forms	anode dissolves
C	colourless gas forms	pink solid forms
D	pink solid forms	colourless gas forms

3. The heat-reflecting shields of space rockets are electroplated with gold. Select the most likely electrodes and electrolyte used to gold-plate the heat shield.

	Negative electrode	Positive electrode	Electrolyte
A	carbon	heat shield	gold compound
B	heat shield	gold	gold compound
C	gold	heat shield	platinum
D	heat shield	carbon	platinum

4. The diagram below shows a simple cell.

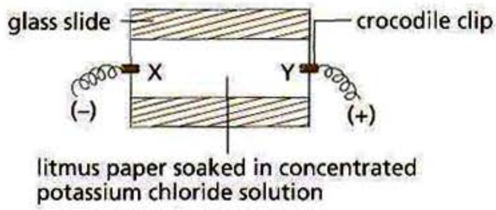


What are metals *X* and *Y*?

	X	Y
A	copper	lead
B	lead	zinc
C	lead	iron
D	zinc	iron

Challenge

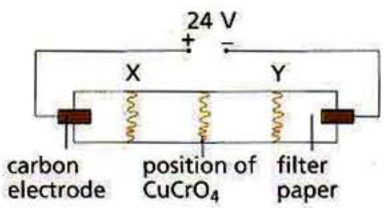
1. The apparatus shown below is connected to a battery. Electric current is passed through the apparatus.



After a few minutes, what colours would be observed on the litmus paper at points X and Y?

	X	Y
A	red	blue
B	red	bleached
C	blue	bleached
D	blue	red

2. a) An apparatus was set up as shown in the diagram below.

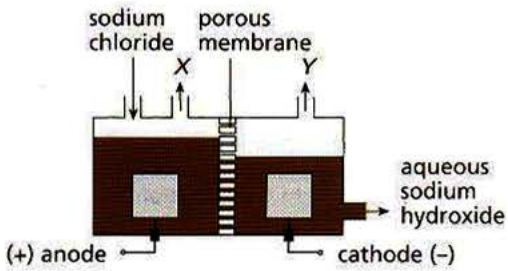


Aqueous copper(II) chromate, $CuCrO_4$, is green due to the blue Cu^{2+} ions and yellow CrO_4^{2-} ions. A drop of concentrated copper(II) chromate was placed on a strip of filter paper moistened with distilled water. The filter paper was then connected to a 24 V battery through carbon electrodes. After a few minutes, two coloured bands were formed at X and Y.

i) Does the apparatus demonstrate a simple cell or an electrolytic cell? Explain your answer.
ii) What is the purpose of making the filter paper moist?
iii) State the colours of the bands at X and Y and explain how they are formed.

b) With reference to molten lead and molten lead(II) bromide, state two differences between a metallic conductor and an electrolytic conductor.

3. The diagram below shows the electrolytic cell used commercially to produce sodium hydroxide from concentrated sodium chloride solution.



a) Identify the gases *X* and *Y* which are liberated.

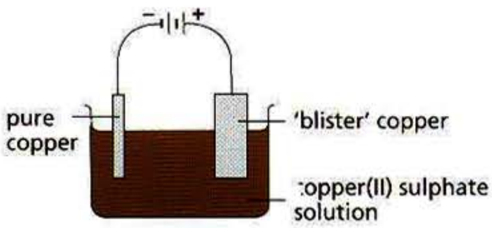
b) Give the equations for the reactions at the electrodes which produce gases *X* and *Y*.

c) Suggest a reason for the presence of a porous membrane.

d) The anode used for this cell is titanium. From this information, suggest a property of titanium.

e) If an iron anode is used, the amount of *X* collected is much reduced. Explain this observation.

4. A block of 'blister' copper contained an impurity. The impurity was a metal that is below copper in the reactivity series. The diagram below shows the apparatus used for refining the copper.



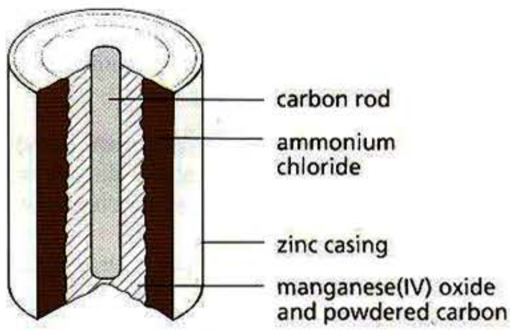
a) Give the formulae of the ions present in aqueous copper(II) sulphate.

b) Write the equation for the reaction occurring at the cathode.

c) The loss of mass of the anode was 52 g and the gain in mass of the cathode was 48 g.
   i) Explain why the anode decreased in mass.
   ii) What is the percentage purity of the 'blister' copper?

d) Explain what you would observe just below the anode after electrolysis?

5. The diagram below shows the cross-section of a dry cell.



When the dry cell is working, the following reactions take place:

At the zinc electrode,

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$

At the carbon electrode,

$$2NH_4^+ + 2e^- \longrightarrow 2NH_3 + H_2$$

a) Identify the positive and negative electrodes in the dry cell. Explain your answer.

b) What is the electrolyte used in the dry cell?

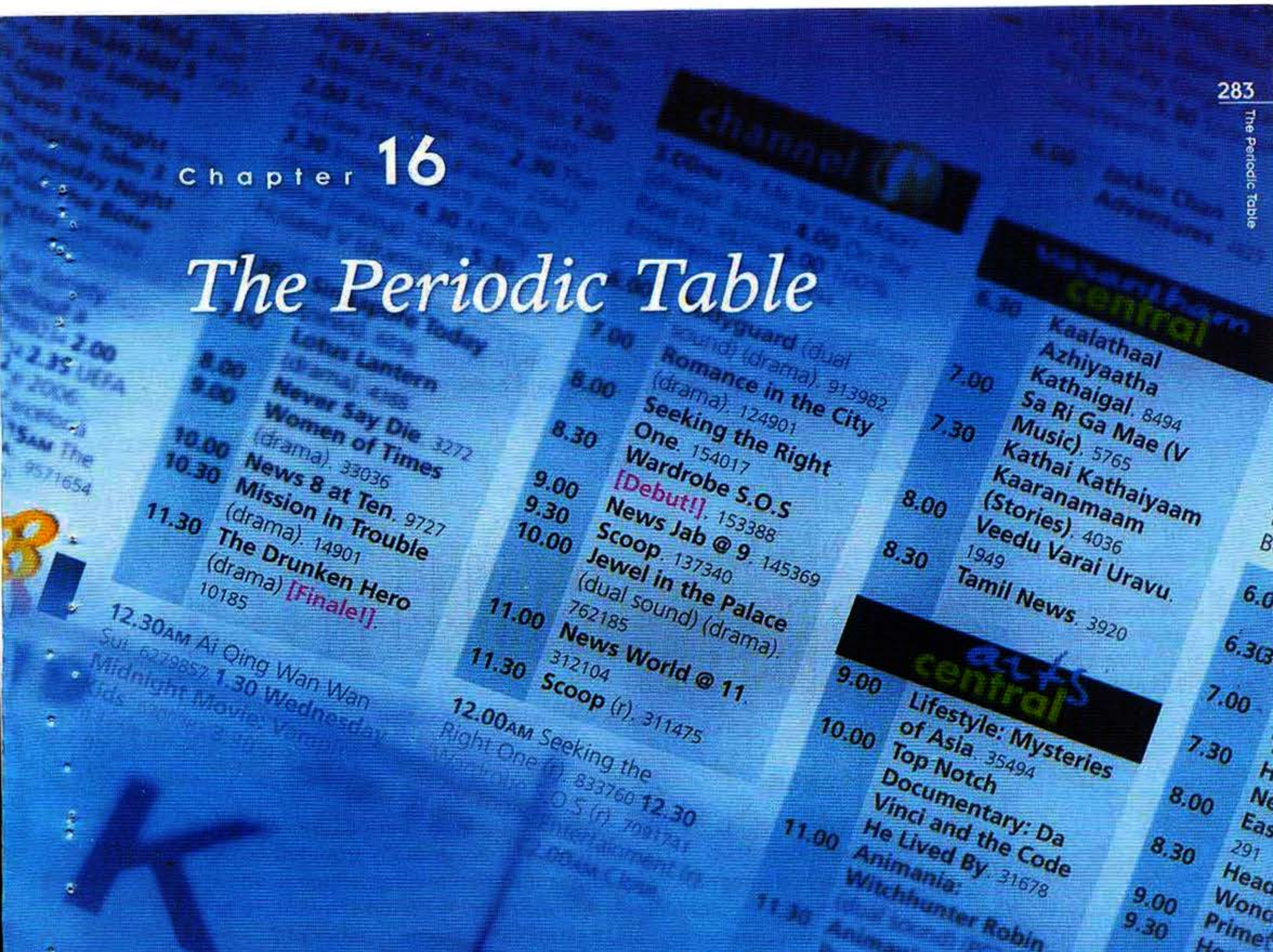
c) What type of particles must be present in a solution which conducts electricity?

d) Why must the paste of the electrolyte in a dry cell not be completely dry?

e) What will happen to the zinc casing when current is taken from the cell?

Chapter 16

The Periodic Table



A table is a very efficient way of displaying a lot of information. At school, you have a timetable to tell you what lessons will be taught at different times of the day, in which rooms, and how long each lesson will last.

You probably watch television at home. You can use the TV listing, which is also a table, to tell at a glance which programmes are on. You can then select the programme you want to watch.

Chemists have a table to help them organise information too. It is called the **Periodic Table**. You have learnt in chapter 4 that the elements and their symbols are recorded in the Periodic Table. The Periodic Table is useful to chemists because it can be used to predict the properties of an element based on its position in the table.

Chapter Outline

Dmitri Mendeleev (1834 – 1907)

The Russian chemist, Mendeleev, is remembered for organising the elements into the Periodic Table. He was a Russian patriot and worked hard for his country. He was one of the first Russian professors to write a textbook for his chemistry students. He received many honours, the greatest of which is having an element named after him. Element 101 is Mendelevium.

16.1 Features of the Periodic Table

The Periodic Table is a list of elements arranged in order of their *increasing proton (atomic) numbers.* The Periodic Table divides the elements into periods and groups. A **period** is *a horizontal row of elements* and a **group** is *a vertical column of elements.*

Fig. 16.1 shows the Periodic Table. Study the table carefully and you will notice several important features about it.

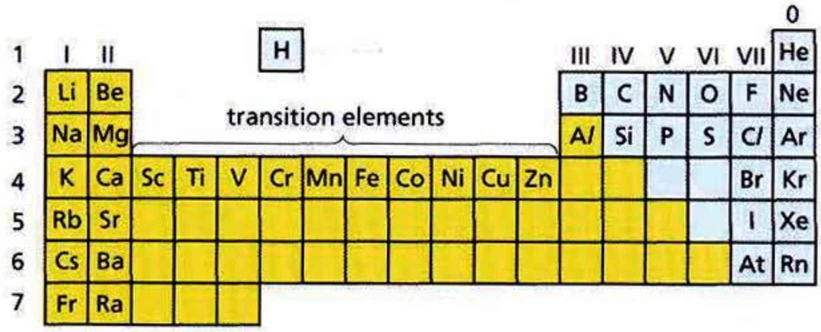


Fig. 16.1 A simplified Periodic Table

Periods

The Periodic Table consists of seven periods of elements, numbered 1 to 7. The periods run horizontally from left to right. Each element in a period has a proton number which is one less than the element after it. For example, in Period 1, hydrogen has a proton number of 1 and helium has a proton number of 2.

Groups

The Periodic Table has eight groups of elements, numbered from I to 0. They run vertically from top to bottom. Group 0 is sometimes called Group VIII.

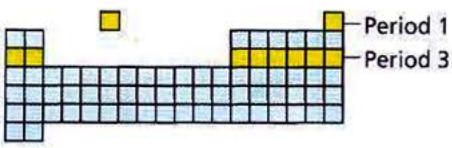


Fig.16.2 The positions of Period 1 and Period 3 in the Periodic Table

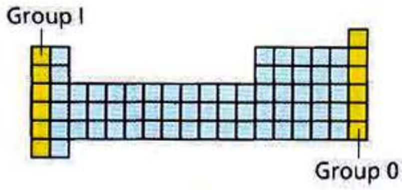


Fig.16.3 The positions of Group I and Group 0 in the Periodic Table

Transition Elements

The block of metals between Groups II and III is known as the **transition elements**.

16.2 | Periodic Trends

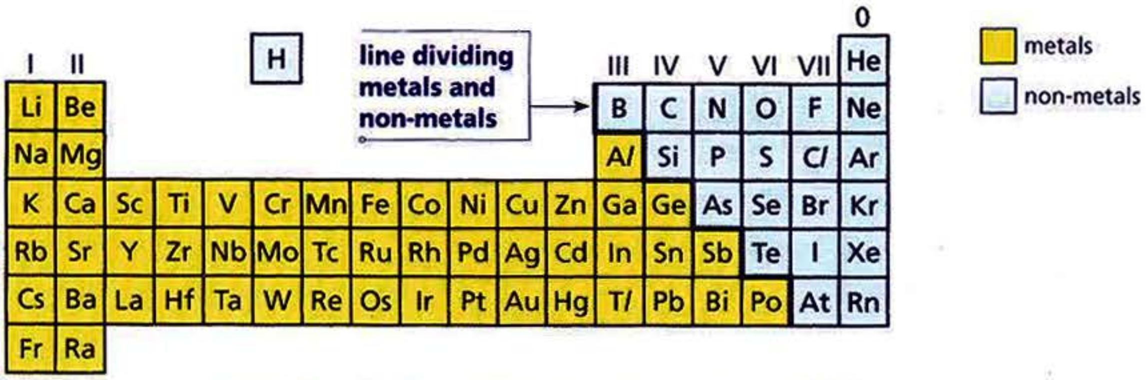


Fig. 16.4 The positions of metals and non-metals in the Periodic Table

The Periodic Table divides the elements into metals and non-metals. The bold line in Fig. 16.4 divides the metals from the non-metals.

Several elements, for example silicon (Si) and germanium (Ge), are located close to the bold line. Because of their positions, these elements have the properties of both a metal and non-metal. Silicon and germanium are known as **metalloids**.

TidBit

Metalloids are used to make computer chips.

Metallic and Non-metallic Characteristics

Metals are grouped on the left-hand side of each period. Non-metals are grouped on the right-hand side.

Due to the change from metal to non-metal across a period, there is also a change in the properties of the elements (Table 16.1).

Group	I	II	III	IV	V	VI	VII	0
Symbol	Na	Mg	Al	Si	P	S	Cl	Ar
Name	sodium	magnesium	aluminium	silicon	phosphorus	sulphur	chlorine	argon
Properties	← metallic →			metalloid	← non-metallic →			
Nature of oxide	← basic →		amphoteric	← acidic →				

Table 16.1 Properties of elements across Period 3

From left to right across a period, there is a decrease in metallic properties and an increase in non-metallic properties.

What do group and period numbers tell us about the electronic structures of elements?

The electronic configurations of Group II, IV and VI elements are given in Table 16.2.

Element	Beryllium	Magnesium	Calcium	Carbon	Silicon	Oxygen	Sulphur
Electronic configuration	2, 2	2, 8, 2	2, 8, 8, 2	2, 4	2, 8, 4	2, 6	2, 8, 6
Group	II	II	II	IV	IV	VI	VI
Period	2	3	4	2	3	2	3

Table 16.2 Some elements of Groups II, IV and VI, and their electronic configurations

The table shows that Group II elements have two valence electrons, Group IV elements have four valence electrons and Group VI elements have six valence electrons. Thus, *the number of valence electrons (outer shell electrons) is the same as the group number of the element.*

Furthermore, since elements with similar electronic configurations have similar chemical properties, we can deduce that *elements in the same group have similar chemical properties.*

Beryllium, carbon and oxygen belong to Period 2. They each have two electron shells. Magnesium, silicon and sulphur belong to Period 3 and they each have three electron shells. Thus, *the number of electron shells is the same as the period number of the element.*

TidBit

Transition elements are the block of elements between Groups II and III of the Periodic Table. They are so called because they represent the transition from metals to non-metals as we move from left to right across a period.

Transition elements are also known as transition metals. They can exist in more than one oxidation state, e.g. iron can form the ions Fe^{2+} and Fe^{3+}.

Chem-Aid

When an element has a tendency to lose electrons, we say that it is metallic.

What are the trends when going down a group?

In any group of the Periodic Table, the following trends can be seen as you go down the group:

- The proton number becomes bigger.
- The atoms become bigger.
- The properties of the elements become more metallic. This is because on going down the group, elements lose electrons more easily.

How are the types of ion formed by elements related to group number?

Table 16.3 shows the relationship between group number and the charge of the ions formed by each element.

Group number	I	II	III	IV, V	VI	VII	0
Type of ions formed	positive	positive	positive	elements tend to form covalent compounds	negative	negative	elements do not form compounds
Charge of ion	+1	+2	+3		−2	−1	
Examples	Na^+, K^+	Mg^{2+}, Ca^{2+}	Al^{3+}	CH_4, PCl_5	O^{2-}, S^{2-}	F^-, Cl^-	

Table 16.3 **The relationship between group number and the ion formed by an element**

The elements in Groups I, II and III are metals.

- Their atoms *lose electrons to form positive ions,* e.g. Na^+, Mg^{2+} and Al^{3+}.
- The charge of the ion is the same as the group number of the element forming it.

The elements in Groups IV and V of the Periodic Table are less likely to form ions. They

- *share electrons to form covalent bonds,*
- have a maximum oxidation state that is the same as the group number of the element (Table 16.4).

Element	Group number	Oxidation state
carbon	IV	+4 in CH_4
phosphorus	V	+3 in PCl_3 +5 (maximum in PCl_5)

Table 16.4 **The relationship between group number and oxidation state**

The elements in Groups VI and VII

- are non-metals,
- tend to *gain electrons and form negative ions,* e.g. S^{2-} and Cl^-.

The elements in Group 0 of the Periodic Table

- have stable electronic configurations,
- do not form compounds.

The changes from non-metallic to metallic character when going down a group are less obvious for elements in Groups I, II, VII and 0. However, for elements in the middle of the Periodic Table (e.g. Groups IV, V and VI), the changes are significant. For example, in Group IV, the element at the top (carbon) is a non-metal but the elements at the bottom (tin and lead) are metals.

Key ideas

1. The elements in the Periodic Table are arranged in order of increasing proton (atomic) number.
2. The horizontal rows of elements in the Periodic Table are called periods.
3. The vertical columns of elements in the Periodic Table are called groups.
4. As a period is crossed from left to right, the metallic character of the elements decreases while the non-metallic character increases.
5. Elements in the same group have
   - the same number of valence electrons,
   - similar chemical properties.
6. On descending a group, the elements become more metallic.
7. Elements in Groups I to III form positive ions, e.g. Na^+, Mg^{2+}, Al^{3+}.
8. Elements in Group VI and VII form negative ions, e.g. O^{2-}, F^-.

Test Yourself 16.1

Worked Example

An element *X* is located in Period 3, Group III of the Periodic Table. Deduce its electronic configuration.

A (2, 8, 2) B (2, 8, 3) C (2, 2, 8) D (2, 8, 8)

Thought Process

X is found in Period 3, hence it has three electron shells. Since *X* is in Group III, it has three valence electrons. The electronic configuration of *X* is therefore (2, 8, 3).

Answer

B

Question

Using only the elements shown in this Periodic Table, answer the following questions.

			H				
							He
			C				
			Si			Cl	

a) Deduce the electronic structure of chlorine. Explain your answer.

b) Give the symbol of
   i) **two** elements in the same group.
   ii) **two** elements in the same period.
   iii) **two** elements that combine together to form an acid.
   iv) the element whose atoms contain the largest number of electrons.
   v) an element that has an oxidation state of –1.

16.3 | Group I Elements — Alkali Metals

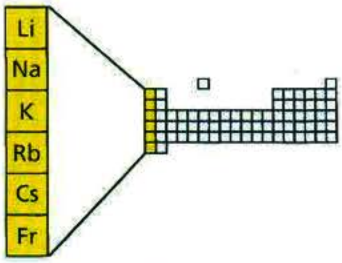
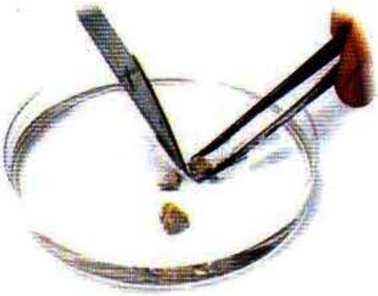


Fig. 16.5 Group I elements — alkali metals

The first three Group I elements are lithium (Li), sodium (Na) and potassium (K). They are called **alkali metals** because they react with water to form alkalis.

Since lithium, sodium and potassium belong to the same group in the Periodic Table, they must have similar physical and chemical properties.

Physical Properties of the Alkali Metals



Alkali metals can be easily cut.

Alkali metals are soft. They can be cut easily. When freshly cut, each of these elements shows a shiny and silvery surface that rapidly tarnishes in air.

The alkali metals have
- low melting and boiling points,
- low densities. Lithium, sodium and potassium float on water.

How do the physical properties change on going down Group I?

Examine Table 16.5. It shows that the melting points and boiling points of the alkali metals decrease on going down the group. The density of the alkali metals generally increases down the group.

Element	Melting point (°C)	Boiling point (°C)	Density (g/cm^3)
lithium	180	1330	0.53
sodium	98	890	0.97
potassium	64	760	0.86
rubidium	38	688	1.53

Table 16.5 Physical properties of some Group I elements

Lithium is used to make batteries.

How can we use the Periodic Table to predict the properties of alkali metals?

Referring to Table 16.5, we can see that the melting points of alkali metals decrease down the group. Hence, the melting point of caesium, the element below rubidium, should be below 38 °C. In fact, it is 28 °C.

Chemical Properties of the Alkali Metals

The alkali metals are reactive metals. They are stored in oil to prevent them from reacting with air and water.

Each alkali metal has one valence electron in the valence shell. By losing this valence electron, an alkali metal attains a noble gas configuration. Due to their similar electronic structures, all alkali metals have similar chemical properties.

How do alkali metals react with water?

All the alkali metals react with cold water to form hydrogen and an alkali. If a drop of the solution is placed on a piece of red litmus paper after the reaction, the litmus paper will turn blue. Table 16.6 describes the reactions of alkali metals with water.

Alkali metal	Observations and equations for reactions with water
lithium	Reacts quickly with water. Lithium floats on the water. No flame is seen. lithium + water $\longrightarrow$ lithium hydroxide + hydrogen $2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$
sodium	Reacts very quickly. Sodium melts. The molten sodium darts around the water surface. A yellow flame is seen. sodium + water $\longrightarrow$ sodium hydroxide + hydrogen $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$
potassium	Reacts violently. Potassium melts. A lilac flame is seen. potassium + water $\longrightarrow$ potassium hydroxide + hydrogen $2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$

Table 16.6 Reactions of alkali metals with water

Going down Group I, alkali metals are more reactive.

Why are alkali metals powerful reducing agents?

All Group I elements form ions with a charge of +1 by losing one electron from the outer shell. For example,

$$Li \longrightarrow Li^+ + e^-$$

Since the alkali metals give away their electrons readily, they behave as powerful reducing agents in all their reactions.

Do all alkali metals have the same reactivity?

As we go down Group I, the size of the atom increases. Sodium is bigger than lithium, and potassium is bigger than sodium. It is easier to lose the valence electron from bigger atoms. Hence, reactivity **increases** on going down Group I. This is illustrated by the reactions of the alkali metals with water (Table 16.6).

Properties of Compounds of the Alkali Metals

Compounds of the alkali metals are ionic, soluble in water, and have similar chemical formulas (Table 16.7).

Element	Carbonate	Nitrate	Sulphate
Li	Li_2CO_3	$LiNO_3$	Li_2SO_4
Na	Na_2CO_3	$NaNO_3$	Na_2SO_4
K	K_2CO_3	KNO_3	K_2SO_4

Table 16.7 Compounds of alkali metals

Key ideas

Group I Elements (Alkali metals)

Electronic structure
- One valence electron
- e.g. Na (2, 8, 1) K (2, 8, 8, 1)

Physical properties
- Soft metals
- Low melting and boiling points
- Low densities
- Compounds are soluble in water.

Chemical properties
- React with water to form an alkali and hydrogen
- Powerful reducing agents
- Form ionic compounds

Group trends

On going down the group, these increase:
- Density
- Reactivity
- Reducing power

However, melting and boiling points decrease.

Test Yourself 16.2

Worked Example

Which statement about a Group I metal *M* is true?

A It reacts with oxygen to form an acidic oxide.
B It will displace hydrogen from cold water.
C Its chloride is insoluble in water.
D The formula of its nitrate will be $M(NO_3)_2$.

Thought Process

All Group I metals react with cold water to form hydrogen and an alkali. The oxides of these metals are basic and their chlorides are soluble in water. The valency of each Group I element is one, therefore the formula of its nitrate is MNO_3.

Answer

B

Questions

1. Name the element which is in the same period as chlorine but in Group I.

2. Rubidium (Rb) is a Group I element. Write the names and formulae of the products formed when rubidium reacts with
   a) water.
   b) chlorine.

3. Lithium, sodium and rubidium are Group I metals. Which of these metals
   a) has the lowest melting point?
   b) is the least reactive?

16.4 | Group VII Elements — Halogens

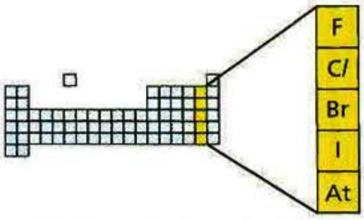


Fig. 16.6 Group VII — halogens

The elements in Group VII of the Periodic Table are called halogens (Fig. 16.6). The elements in the group are fluorine (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (At).

Physical Properties of the Halogens

The halogens are non-metals. They exist as diatomic covalent molecules: F_2, Cl_2, Br_2, I_2. The halogens have low melting and boiling points. They are also coloured. Table 16.8 gives a summary of the physical properties of chlorine, bromine and iodine.

Element	Melting point (°C)	Boiling point (°C)	Appearance
chlorine	–101	–35	greenish-yellow gas
bromine	–7	59	reddish-brown liquid
iodine	114	184	purplish-black solid

Table 16.8 The physical properties of some Group VII elements

Chlorine — a greenish-yellow gas

Iodine — a purplish-black solid

How do the physical properties of the halogens change on going down the group?

On going down the group,

- the melting points and boiling points of the halogens increase. Chlorine is a gas, bromine is a liquid while iodine is a solid at room temperature.
- the colours of the halogens become darker.

How can we use the Periodic Table to predict the properties of halogens?

If we examine Table 16.8, we will notice that the melting points of the halogens increase down the group. Thus, we expect the next halogen to have a melting point greater than 114 °C. Astatine is a solid which melts at 300 °C.

We can also deduce that astatine is black since colour intensity increases down the group.

Chem-Aid

Astatine, like the other halogens, is diatomic and exists as At_2 molecules.

Quick Check

Why are halogens diatomic?

Chemical Properties of the Halogens

Halogens are reactive non-metals. Why are they so reactive? The valence shell of each halogen contains seven valence electrons. This means that only one more electron is needed to achieve a stable noble gas structure.

Halogens react with most metals to form salts called **halides**. Fluoride ions (F^-), chloride ions (Cl^-), bromide ions (Br^-) and iodide ions (I^-) are examples of **halide ions**.

What are the displacement reactions of the halogens?

A **displacement reaction** is a reaction in which one element takes the place of another element in a compound. *A more reactive halogen will displace a less reactive halogen from its halide solution.* For example, when chlorine water is added to aqueous sodium bromide or potassium bromide, a reddish-brown solution is obtained. Chlorine, being more reactive than bromine, displaces bromine from the bromide solution.

Chlorine (Cl_2) is a ***halogen***. Sodium bromide (NaBr) is a ***halide***.

chlorine + sodium bromide ⟶ sodium chloride + bromine
(colourless) (colourless) (colourless) (reddish-brown)

$$Cl_2(aq) + 2NaBr(aq) \longrightarrow 2NaCl(aq) + Br_2(aq)$$

Chlorine also displaces iodine from an iodide solution.

chlorine + potassium iodide ⟶ potassium chloride + iodine
(colourless) (colourless) (colourless) (brown)

$$Cl_2(aq) + 2KI(aq) \longrightarrow 2KCl(aq) + I_2(aq)$$

Similarly, bromine displaces iodine from an iodide solution.

bromine + potassium iodide ⟶ potassium bromide + iodine
(reddish-brown) (colourless) (colourless) (brown)

$$Br_2(aq) + 2KI(aq) \longrightarrow 2KBr(aq) + I_2(aq)$$

However, when bromine water is added to potassium chloride solution, no reaction occurs. This is because bromine is less reactive than chlorine, hence bromine cannot displace chlorine from a chloride solution. *A less reactive halogen cannot displace a more reactive halogen from its halide solution.*

We can deduce the order of reactivity of the halogens from their displacement reactions. Unlike the alkali metals, the reactivity of the halogens **decreases** down the group. Fluorine is the most reactive halogen. Astatine is the least reactive.

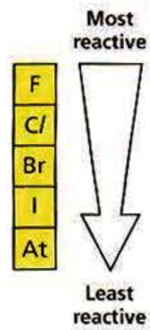


Fig.16.7 The reactivity of halogens decreases down the group

Why are halogens powerful oxidising agents?

During chemical reactions, halogen atoms gain electrons to form halide ions. The halogens are therefore powerful oxidising agents.

$$X_2 + 2e^- \longrightarrow 2X^- \text{ (where } X_2 = \text{halogen; } X^- = \text{halide ion)}$$

Try it Out

Bleach is commonly used to remove stains from clothing. Chloride ions are present in bleach. Is bleaching a redox reaction? You may use the Internet to find out.

The displacement reactions between halogens and other halide ions can also be classified as **redox reactions**. For example, in the displacement reaction between chlorine and potassium bromide, chlorine acts as the *oxidising agent* while the bromide ion acts as the *reducing agent*. Chlorine oxidises bromide ions to bromine and is itself reduced to chloride ions.

$Cl_2(aq) + 2e^- \longrightarrow 2Cl^-(aq)$ (reduction)

$2Br^-(aq) \longrightarrow Br_2(aq) + 2e^-$ (oxidation)

The overall ionic equation for the reaction is

reduced

$$Cl_2(aq) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$$

oxidised

Key ideas

Group VII Elements (Halogens)
General formula: X_2

Electronic structure
- Seven valence electrons
- e.g. F (2, 7) Cl (2, 8, 7)

Physical properties
- They are non-metals.

At room temperature,
- fluorine and chlorine are gases,
- bromine is a liquid,
- iodine is a solid.
- They have low melting and boiling points.

Chemical properties

A more reactive halogen displaces a less reactive halogen from its halide solution.

Group trends

On going down the group, the
- colour becomes darker,
- boiling and melting points increase,
- reactivity decreases,
- oxidising power decreases.

Test Yourself 16.3

Worked Example

Iodine is below bromine in Group VII of the Periodic Table.

a) State **two** physical properties of iodine. Explain your answer.

b) Describe and explain what would happen when bromine water is added to potassium iodide solution.

Answer

a) Iodine is a black solid with low melting and boiling points. Since chlorine is a greenish-yellow gas and bromine is a reddish-brown liquid at room temperature, we expect iodine to be a black solid.

Since iodine is a solid, its melting point and boiling point would be above room temperature. However, since it is a non-metal, its melting and boiling points are low.

b) The potassium iodide solution would turn from colourless to brown. Bromine is more reactive than iodine so it will displace iodine from potassium iodide. Bromine will be reduced to bromide ions and the iodide ions will be oxidised to iodine.

$$Br_2(aq) + 2KI(aq) \longrightarrow I_2(aq) + 2KBr(aq)$$

Question

Astatine, At, is the element below iodine in the Periodic Table.

a) Suggest the formula of the compound formed between magnesium and astatine.
b) What is the name of this compound?
c) Predict a physical property of this compound.
d) Predict the chemical changes that take place, if any, when
   i) chlorine is bubbled into an aqueous solution of NaAt.
   ii) astatine is added to aqueous potassium iodide.

16.5 | Group 0 Elements — Noble Gases

The elements in Group 0 or Group VIII are called **noble gases** (Fig. 16.8). They are the least reactive elements in the Periodic Table. Apart from helium, the other noble gases have eight valence electrons. Helium has two valence electrons. Their full electronic structures make the noble gases unreactive.

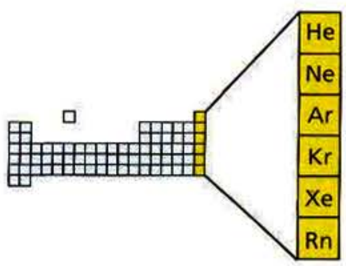


Fig. 16.8 Group 0 — noble gases

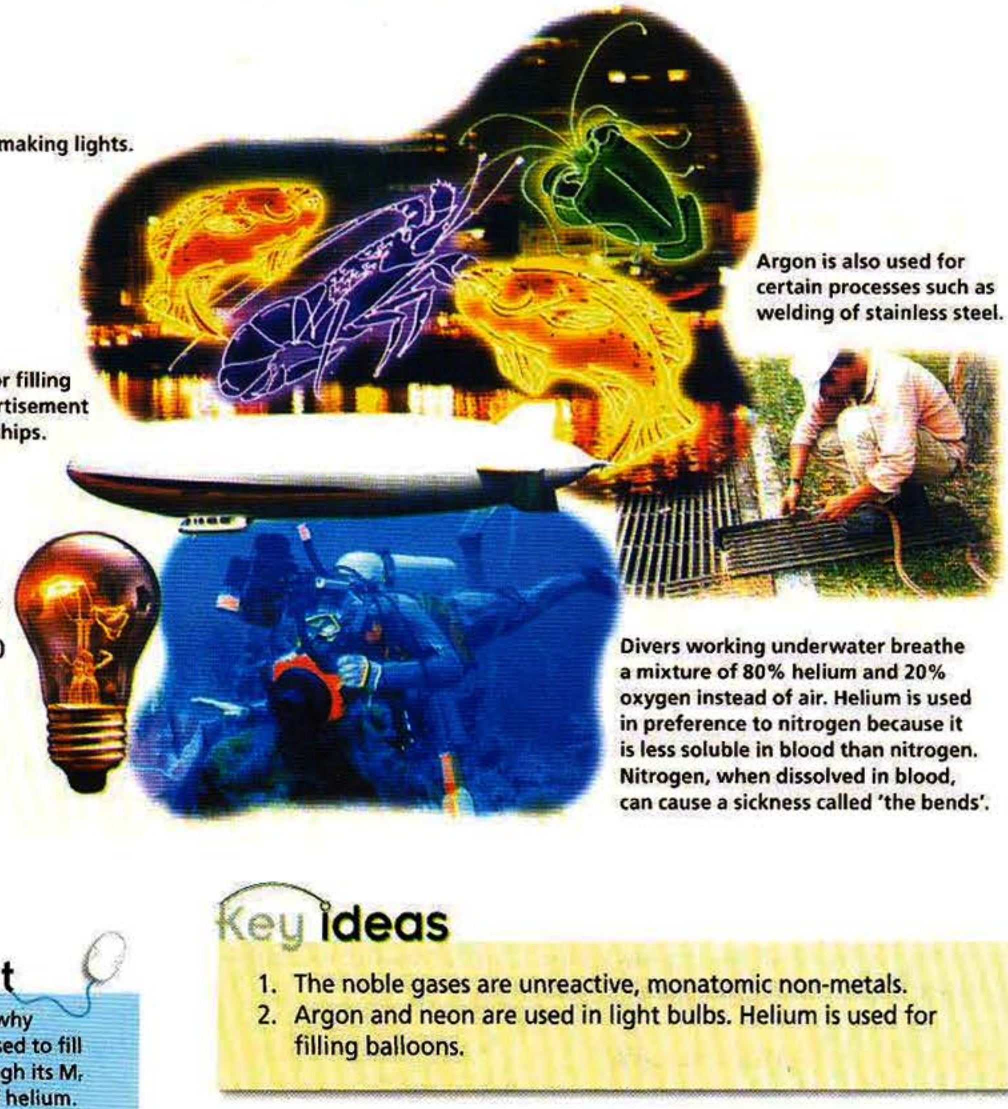
Properties of Noble Gases

The Group 0 elements are also referred to as **inert gases** (because they are unreactive) or **rare gases** (because less than 1% of the air is made up of these gases).

The noble gases

- are monatomic elements,
- are all colourless gases at room temperature,
- have low melting and boiling points that increase on going down the group,
- are insoluble in water,
- are unreactive.

Noble gases do not react to form compounds because their atoms have full outer shells of electrons. *They do not lose, gain or share electrons, hence they are unreactive*. Noble gases such as argon are used to fill light bulbs. They provide an inert atmosphere which prevents the filament from oxidation.



Neon is used in making lights.

Argon is also used for certain processes such as welding of stainless steel.

Helium is used for filling weather or advertisement balloons and airships.

Argon is used to fill electric bulbs. It provides an inert (unreactive) atmosphere to help protect the filament from oxidation in air.

Divers working underwater breathe a mixture of 80% helium and 20% oxygen instead of air. Helium is used in preference to nitrogen because it is less soluble in blood than nitrogen. Nitrogen, when dissolved in blood, can cause a sickness called 'the bends'.

Try it Out

You may wonder why hydrogen is not used to fill airships even though its M_r is less than that of helium. Find out from the Internet by keying in 'Hindenburg' into an Internet search engine.

Key ideas

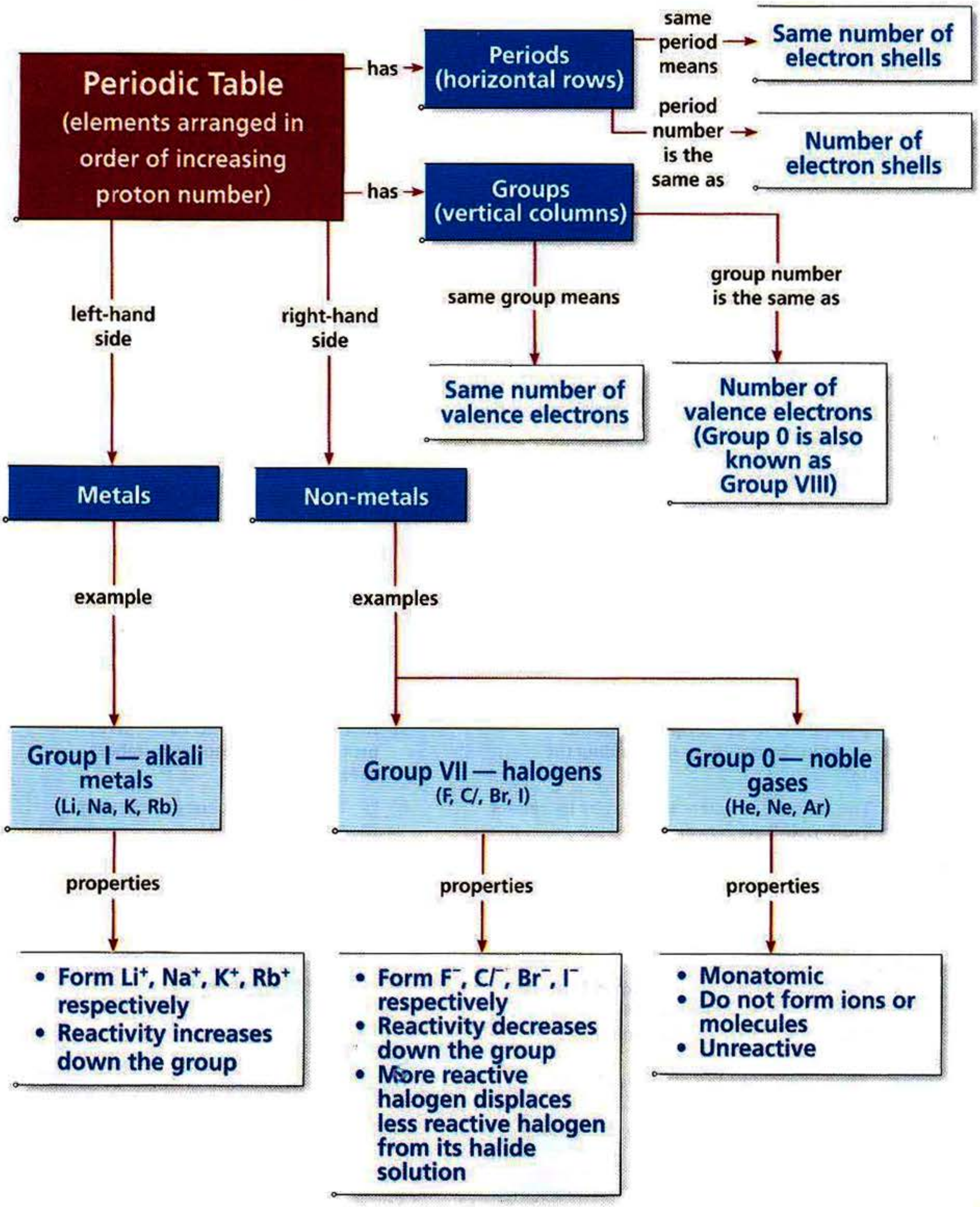
1. The noble gases are unreactive, monatomic non-metals.
2. Argon and neon are used in light bulbs. Helium is used for filling balloons.

Test Yourself 16.4

Questions

1. Explain why noble gases such as helium and argon are used in welding.
2. Rocket fuels are highly explosive and need to be kept in an inert atmosphere. Suggest why helium is used to provide such an atmosphere instead of nitrogen.

Concept Map


Periodic Table
(elements arranged in order of increasing proton number)
has
Periods (horizontal rows)
same period means
Same number of electron shells
period number is the same as
Number of electron shells
has
Groups (vertical columns)
same group means
Same number of valence electrons
group number is the same as
Number of valence electrons (Group 0 is also known as Group VIII)
left-hand side
right-hand side
Metals
Non-metals
example
examples
Group I — alkali metals
(Li, Na, K, Rb)
Group VII — halogens
(F, Cl, Br, I)
Group 0 — noble gases
(He, Ne, Ar)
properties
properties
properties
• Form Li^+, Na^+, K^+, Rb^+ respectively
• Reactivity increases down the group
• Form F^-, Cl^-, Br^-, I^- respectively
• Reactivity decreases down the group
• More reactive halogen displaces less reactive halogen from its halide solution
• Monatomic
• Do not form ions or molecules
• Unreactive

Exercise 16

Foundation

1. Which statement about the Periodic Table is **false**?
   - A Elements get more metallic on descending a group.
   - B Elements in the same group have the same number of valence electrons.
   - C From left to right across a period, there is a change from non-metallic to metallic character.
   - D The elements are arranged in order of increasing proton number.

2. Why are lithium and fluorine placed in Period 2 of the Periodic Table?
   - A They both form ions with a charge of 2.
   - B They both have two electron shells.
   - C They can react together to form LiF.
   - D They have two electrons in their first shell.

3. Astatine is below iodine in the Periodic Table. Which of the following best describes astatine at room temperature and pressure?
   - A Black solid
   - B Black liquid
   - C Reddish-brown liquid
   - D Reddish-brown solid

4. What happens to the properties of the elements in Group VII on descending the group?
   - A The number of valence electrons increases.
   - B Their oxidising power increases.
   - C They become more metallic.
   - D They decrease in reactivity.

5. Which statement about all the ions of the Group VII elements is correct?
   - A They contain equal numbers of protons and neutrons in their nuclei.
   - B They contain more electrons than protons.
   - C They contain seven electrons in the outer shell.
   - D They have a single positive charge.

6. Which element is in Group 0?
   - A Helium
   - B Iron
   - C Lithium
   - D Nitrogen

7. Which compound consists only of elements which are in the same period of the Periodic Table?
   - A KCN
   - B Li_2SiO_3
   - C $NaAlCl_4$
   - D $SOCl_2$

Challenge

1. An element *M* forms an ion M^-. The electronic configuration of this ion is (2, 8). Which statement is true?
   - A *M* is in Group VI and Period 2.
   - B *M* is in Group VII and Period 2.
   - C *M* is in Group VII and Period 3.
   - D *M* is in Group VI and Period 3.

2. Which statement about the halogens is correct?
   - A Bromine has more electrons than iodine.
   - B Silver chloride is soluble in water.
   - C They form ions by gaining one electron.
   - D They are monatomic.

3. a) Refer to the Periodic Table. Write the proton number (atomic number) of caesium, Cs.

   b) i) Predict the valency of caesium.
   ii) Write the formula for caesium oxide.

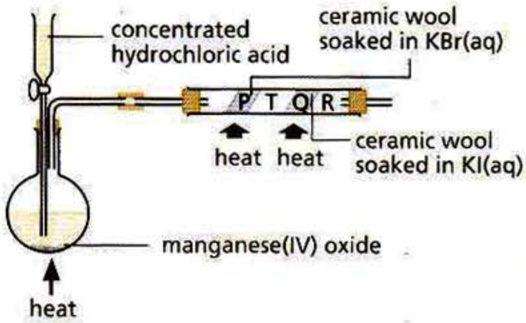
   c) i) Caesium and sodium both react with cold water. Which reacts faster?
   ii) Construct an equation for the reaction of caesium with water. Include state symbols.
   iii) Suggest how caesium is stored. (*C*)

4. a) State **two** properties that are used to arrange elements in the Periodic Table.
   b) Use three elements from Group VII of the Periodic Table to illustrate the trend in
      i) physical property.
      ii) chemical property.
   c) Chlorine dissolves in hot aqueous sodium hydroxide to form sodium chloride and sodium chlorate(V). The following equation represents the reaction.

$$3Cl_2 + 6NaOH \longrightarrow 5NaCl + NaClO_3 + 3H_2O$$

      i) Write an equation to show the reaction that occurs between bromine and sodium hydroxide.
      ii) Describe any colour changes during the reaction.

5.



Chlorine can be made by reacting concentrated hydrochloric acid with manganese(IV) oxide (MnO_2) using the apparatus above. The other products of the reaction are manganese(II) chloride ($MnCl_2$) and water.

a) i) Write the equation for this reaction.
   ii) Why is the reaction in (a)(i) a redox reaction?

b) What is the colour of
   i) potassium bromide and
   ii) potassium iodide solution?

c) State, with reasons, what will be formed at T and R. Write equations for any reactions that occur.

d) Why is this experiment performed in a fume cupboard?

Chemistry Today

The first person to try and classify elements was Lavoisier. In 1789, he divided the then-known 'elements' into four groups:

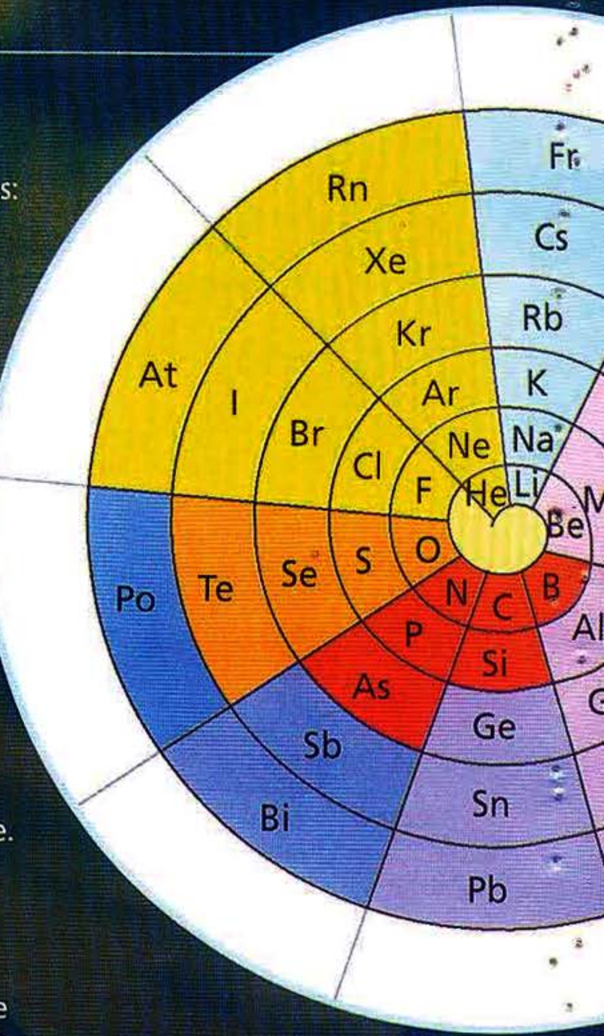
- Simple substances such as oxygen and hydrogen
- Non-metallic substances such as sulphur and phosphorus
- Metallic substances such as silver and mercury
- Earthy substances such as calcium oxide and magnesium oxide

Calcium oxide and magnesium oxide were thought to be elements because the chemists then could not break them down into their elements.

In 1860, John Newlands, an English chemist, tried to arrange the 63 elements that the chemists in his time knew in order of their relative atomic mass. He noticed that the properties of every eighth element were similar.

In 1869, Dmitri Mendeleev produced the modern Periodic Table. He left gaps for elements that he thought had not yet been discovered. He also accurately predicted the properties of the missing elements.

Today, chemists are still exploring new ways to represent all the elements. One way is shown here. It is called the Benfey version.



CRITICAL THINKING

The Periodic Table is an arrangement of the elements according to their proton numbers so that elements with similar properties are in the same column. Suggest why chemists are still trying to find new ways of presenting the Periodic Table.

Enter the words 'Periodic Table' and click on 'Images' in your search engine. You will see a large number of different versions of the table such as those by Stowe and Tarantola. Find out how the elements are arranged in these representations.

Chapter 17

Energy Changes

Chapter Outline

When you have a headache, applying a cold pack can relieve the pain. Some cold packs contain chemicals which take in heat when they react. This is why cold packs have a cooling effect on their surroundings.

In this chapter, you will learn about reactions that take in and give out heat, and their applications in our daily lives.

17.1 | Exothermic and Endothermic Changes

Energy cannot be created or destroyed. However, it can be changed from one form to another. Energy changes occur in chemical reactions and even in some physical processes, such as when a solid dissolves in water.

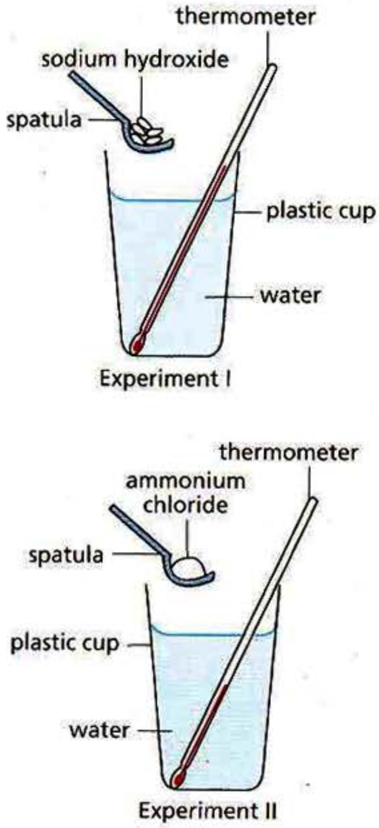


Fig. 17.1 Measuring temperature change during a reaction

In the two experiments shown in Fig. 17.1, the temperature changes are measured when a solid is dissolved in water. In experiment I, sodium hydroxide pellets are added to water. The mixture is carefully stirred to dissolve the pellets. The temperature of the water is recorded before and after adding sodium hydroxide. In experiment II, the experiment is repeated using ammonium chloride crystals instead of sodium hydroxide.

The results of the experiments are recorded in Table 17.1.

Experiment	Solid	Initial temperature of water (°C)	Temperature after adding solid (°C)
I	NaOH	28	34
II	NH_4Cl	28	22

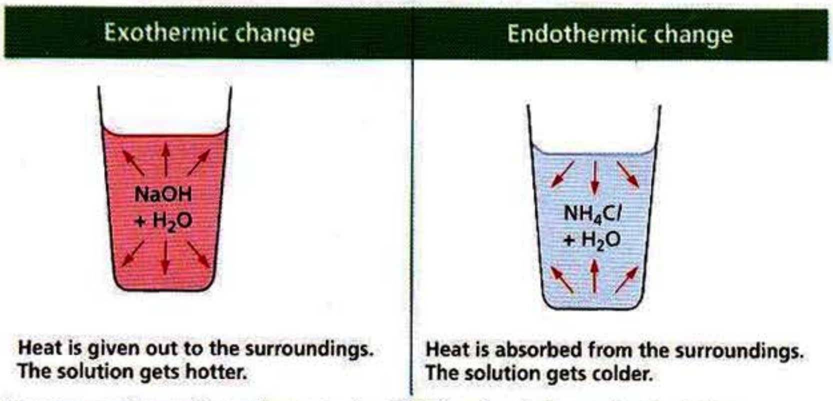
Table 17.1 Results of experiments I and II

What conclusions about energy changes can be made from these two experiments?

In experiment I, heat energy was given out when the solid dissolved in water. Thus, the temperature of the solution rose. We say this change is **exothermic**. In experiment II, heat energy was absorbed from the surroundings when the solid dissolved in water. As a result, the temperature of the solution dropped. We say this change is **endothermic**.

Science Skills

Why are plastic or polystyrene cups used for heat experiments instead of metal containers?



(The surroundings refer to the water in which the chemicals are dissolved, the container that holds the chemicals, the air, etc.)

Fig. 17.2 Exothermic and endothermic changes ('→' indicates direction of heat transfer)

We have just seen that heat energy is given out during a physical process, such as when sodium hydroxide dissolves in water. Heat energy may also be given out during a chemical reaction, such as when sodium hydroxide reacts with hydrochloric acid.

Exothermic Reactions

Reactions that give out heat energy to the surroundings are called **exothermic reactions.**

What are the characteristics of exothermic reactions?
When an exothermic reaction occurs,
a) heat is liberated and is transferred from the chemicals to the surroundings, and
b) the temperature of the reaction mixture rises. The container feels hot.

Fig. 17.3 shows the changes in temperature when an exothermic reaction occurs. Initially, the temperature of the reaction mixture rises until the highest temperature is reached. When the reaction is completed, the temperature of the reaction mixture falls until it reaches room temperature.

Examples of exothermic reactions include
- the combustion of fuels,
- the rusting of iron,
- the corrosion of metals,
- the reaction between acid and alkali (neutralisation),
- respiration.

TidBit

There are several types of hot packs. One type contains a small inner bag of water and a stronger outer bag containing an anhydrous salt. The salt can be anhydrous calcium chloride or anhydrous magnesium sulphate. When the pack is squeezed, the small inner bag breaks and the anhydrous salt dissolves in water. This process is exothermic.

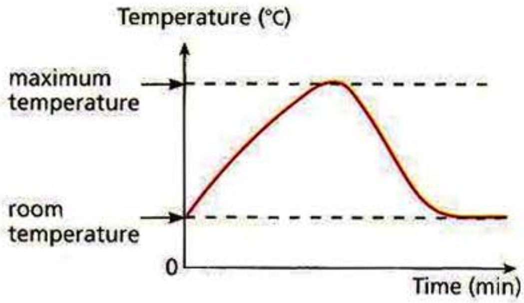


Fig. 17.3 Variation of temperature with time for an exothermic reaction

Endothermic Reactions

Not all reactions give out heat. *Reactions that absorb heat from the surroundings* are called **endothermic reactions.**

What are the characteristics of endothermic reactions?
When an endothermic reaction occurs,
a) heat energy is absorbed and is transferred from the surroundings to the reactants and,
b) the temperature of the reaction mixture falls. The container feels cold.

Fig. 17.4 shows the changes in temperature when an endothermic reaction occurs. Initially, the temperature of the reaction mixture falls until the lowest temperature is reached. When the reaction is completed, the temperature of the reaction mixture rises until it reaches room temperature.

Examples of endothermic reactions include
- photosynthesis,
- the action of light on silver bromide in photographic film,
- thermal decomposition. For example,

calcium carbonate $\xrightarrow{\text{heat}}$ calcium oxide + carbon dioxide

$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$

Chem-Aid

Physical processes can also be classified as exothermic or endothermic.
- Condensation, freezing and dissolving of acids (especially concentrated acids) in water, are exothermic.
- Evaporation, melting, and the dissolving of some ionic compounds (e.g. ammonium chloride, potassium nitrate and copper(II) sulphate crystals) in water are endothermic.

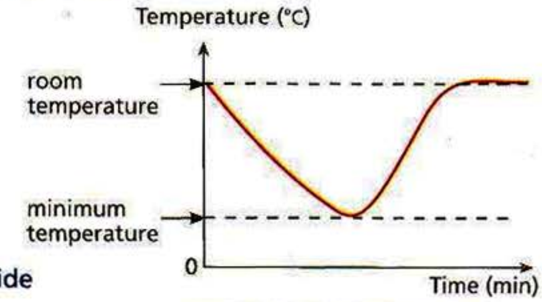


Fig. 17.4 Variation of temperature with time for an endothermic reaction

James Prescott Joule (1818 – 1889)

Scientists named the SI unit for energy as 'Joule' in honour of James Prescott Joule, a British scientist. He discovered how the amount of work done is related to the amount of heat produced. This discovery influenced many theories, including the kinetic particle theory.

Chem-Aid

1. The heat change or enthalpy change of a reaction is also called the **heat of reaction.**
2. A chemical equation that includes the heat change at the right side of the equation is called a **thermochemical equation.**

Heat Changes in a Reaction

The amount of energy involved in a reaction is known as the **heat change** or **enthalpy change** of the reaction. It is measured in **kilojoules (kJ)** and represented by the symbol ΔH. Δ is the Greek letter 'delta', which means change. H means energy content.

For an *exothermic reaction*, *ΔH is negative*. This is because the chemicals have lost energy to the surroundings. For an *endothermic reaction*, *ΔH is positive* because the chemicals gain energy from the surroundings.

When one mole of carbon is burnt in excess oxygen, 349 kJ of heat is produced. This is an exothermic reaction. We can therefore say that for this reaction, ΔH is –349 kJ/mol.

carbon + oxygen ⟶ carbon dioxide

$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -349 \text{ kJ/mol}$

When one mole of hydrogen reacts with one mole of iodine, 52 kJ of heat is absorbed from the surroundings. This is an example of an endothermic reaction. We say that for this reaction, ΔH is +52 kJ/mol.

hydrogen + iodine ⟶ hydrogen iodide

$H_2(g) + I_2(g) \longrightarrow 2HI(g) \quad \Delta H = +52 \text{ kJ/mol}$

Energy Level Diagrams for Exothermic and Endothermic Reactions

A convenient way to show energy changes in a reaction is by means of **energy level diagrams**. Consider an exothermic reaction. In this type of reaction, heat is given out to the surroundings. This means that the total energy of the products is less than that of the reactants. Fig. 17.5 shows the energy level diagram for an exothermic reaction.

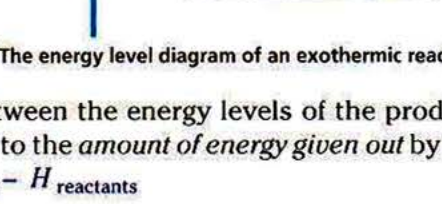


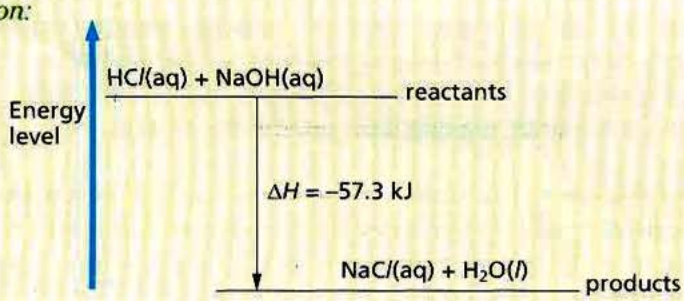
Fig. 17.5 The energy level diagram of an exothermic reaction

The *difference* between the energy levels of the products and the reactants is equal to the *amount of energy given out* by the reaction, i.e. $\Delta H = H_{\text{products}} - H_{\text{reactants}}$

Example 1

When 1 mol of hydrochloric acid reacts with 1 mol of sodium hydroxide, 57.3 kJ of heat is produced. Draw an energy level diagram for this reaction.

Solution:



How does the energy level diagram for an endothermic reaction differ from that of an exothermic reaction?

Since an endothermic reaction absorbs heat from the surroundings, the products will have more energy than the reactants (Fig. 17.6). The *difference* between the energy levels of the products and reactants is equal to the *energy absorbed* during the reaction.

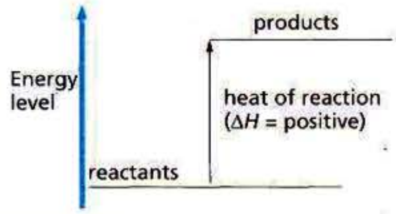


Fig. 17.6 The energy level diagram of an endothermic reaction

Key ideas

1.

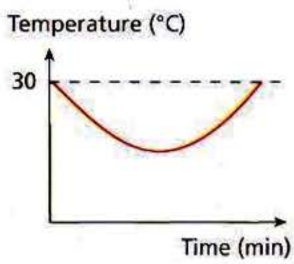
Exothermic reaction	Endothermic reaction
• gives out heat to the surroundings • causes an increase in temperature • has a negative ΔH • has products that have lower energy than the reactants	• takes in heat from the surroundings • causes a decrease in temperature • has a positive ΔH • has products that have higher energy than the reactants

2. An energy level diagram is used to show the heat change (enthalpy change) in an exothermic or endothermic reaction.

Test Yourself 17.1

Questions

1. To decompose 100 g of calcium carbonate into calcium oxide and carbon dioxide, 178 kJ of heat must be supplied. Write the chemical equation, including the ΔH value in kJ/mol, for the reaction.
2. The graph shows how the temperature changes as a substance X is added to water and then the solution is left to stand. Is dissolving X in water an exothermic or endothermic process?
3. When 1 mol of hydrogen reacts with 1 mol of iodine to give 2 mol of hydrogen iodide, 52 kJ of heat is absorbed. Draw an energy level diagram for this reaction.



Magnesium powder is used in fireworks. Where does the light and heat energy come from?

17.2 | Bond Breaking and Bond Making

What causes energy changes in chemical reactions?

All reactions involve either the breaking of bonds or the making of new bonds, or both. For example, when nitrogen reacts with oxygen to form nitrogen monoxide, first the bonds between the nitrogen molecules and oxygen molecules must be broken to form atoms of nitrogen and oxygen. Nitrogen atoms then form bonds with oxygen atoms, making nitrogen monoxide (Fig. 17.7).

Energy changes in reactions are caused by the making and breaking of chemical bonds. When bonds between atoms are broken, heat energy is absorbed. However, when bonds are formed between atoms, heat energy is given out. In other words,

- *bond breaking is an endothermic process.*
- *bond formation is an exothermic process.*

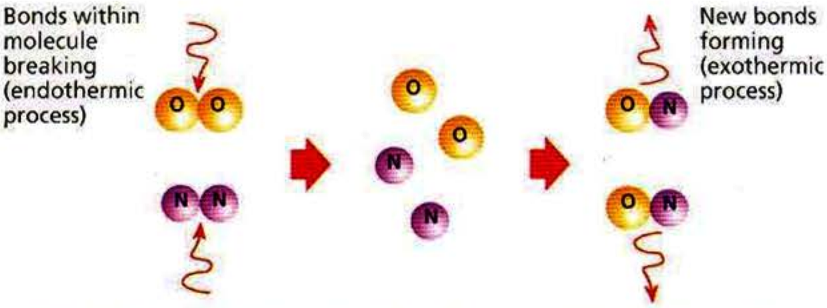


Fig. 17.7 Bond breaking and bond forming processes cause the energy changes in chemical reactions. (' ' indicates the direction of energy transfer.)

How can you tell if a reaction is exothermic or endothermic?

If the energy required for bond breaking is less than the energy given off in bond making, the reaction is exothermic.

Exothermic reaction: $\Delta H_{\text{bond breaking}} < \Delta H_{\text{bond making}}$

If the energy required for bond breaking is more than the energy given off in bond making, the reaction is endothermic.

Endothermic reaction: $\Delta H_{\text{bond breaking}} > \Delta H_{\text{bond making}}$

Let us consider the reaction between hydrogen and oxygen to form water.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

Is this reaction exothermic or endothermic? To find out, study the energy involved in bond breaking and bond forming during the reaction.

Chem-Aid

You learnt in section 17.1 that some reactions such as combustion are always exothermic. This means that in combustion $\Delta H_{\text{bond making}}$ is always greater than $\Delta H_{\text{bond breaking}}$. For endothermic reactions, the reverse is true.

1. Bond breaking

The bonds in the hydrogen molecule and the oxygen molecule must first be broken before the hydrogen atoms and oxygen atoms can combine. This requires energy to be absorbed, i.e. the process is endothermic.

Energy required to break 1 mol of H – H bonds = +436 kJ

$\text{H}-\text{H} \rightarrow 2\text{H} \quad \Delta H = +436 \text{ kJ}$

Energy required to break 2 mol of H – H bonds = 2 × (+436) kJ
= +872 kJ

Therefore, ΔH = +872 kJ

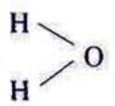
$\text{O}=\text{O} \longrightarrow 2\text{O} \qquad \Delta H = +496 \text{ kJ}$

Energy required to break the bonds in the hydrogen molecules and the oxygen molecule = (+872) + (+496) = +1368 kJ

2. Bond making

The second step involves the joining together of oxygen and hydrogen atoms by forming new O–H bonds. This is an exothermic process.

Energy released on forming 1 mol of O–H bonds = –468 kJ
Energy released on forming 2 mol of O–H bonds in 1 mol of H_2O molecules = 2 × (–468) kJ = –926 kJ

H + H + O ⟶  ΔH = –926 kJ

Therefore, energy released on forming 2 mol of H_2O molecules = 2 × (–926) = –1852 kJ

The overall heat change, or enthalpy change, for the reaction = (+1368) + (–1852) = – 484 kJ

The negative value of the overall ΔH indicates that the reaction is exothermic. We can summarise the processes in the above reaction by using an energy level diagram as shown in Fig. 17.8.

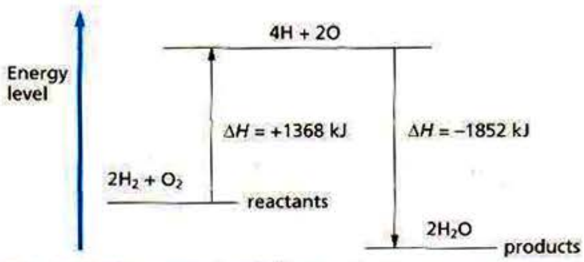


Fig. 17.8 The energy level diagram for the formation of water

Note that *the energy required to break a chemical bond is the same as the energy released when the same chemical bond is formed.*

Thus, H — H $\longrightarrow$ 2H $\quad \Delta H = +436$ kJ
2H $\longrightarrow$ H — H $\quad \Delta H = -436$ kJ

This energy is called the **bond energy** of the H–H covalent bond.

The stronger a bond is, the more energy required to break a bond. Thus, the higher its bond energy.

Let us look at another reaction: the formation of nitrogen dioxide from nitrogen and oxygen. The changes in energy when nitrogen reacts with oxygen to form nitrogen dioxide can be represented by the energy level diagram shown below.

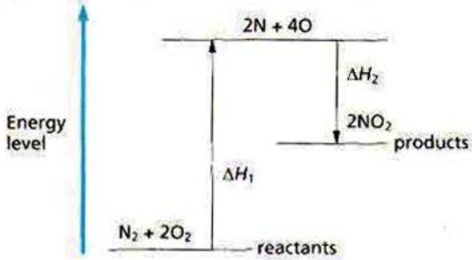


Fig. 17.9 The energy level diagram for the formation of nitrogen dioxide

Key ideas

1. Bond breaking is an endothermic process. Bond making is an exothermic process.
2. A reaction is exothermic if $\Delta H_{\text{bond breaking}} < \Delta H_{\text{bond making}}$
3. A reaction is endothermic if $\Delta H_{\text{bond breaking}} > \Delta H_{\text{bond making}}$

What information can be obtained from this energy level diagram?

The energy level diagram tells us that the formation of nitrogen dioxide from nitrogen and oxygen is an endothermic reaction. It also shows that

- the products have more energy than the reactants,
- the energy required to break the bonds in the nitrogen molecule and the oxygen molecule (ΔH_1) is more than the energy released (ΔH_2) when the nitrogen-oxygen bonds are formed.

Test Yourself 17.2

Worked Example

The bond energy is the energy required to break a covalent bond between two atoms.

Bond	Bond energy (kJ/mol)
H — H	436
H — Cl	431
Cl — Cl	242

Calculate the overall ΔH for the following reaction:

$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

Thought Process

ΔH (reaction) = heat absorbed on bond breaking + heat released on bond breaking

Step 1: Calculate the energy change in breaking 1.0 mol of H — H bonds and 1.0 mol of *Cl* — *Cl* bonds.
Step 2: Calculate the energy change in forming 2.0 mol of H — *Cl* bonds
Step 3: Overall heat change, $\Delta H = \Delta H_{\text{bond breaking}} + \Delta H_{\text{bond making}}$

Answer

$\Delta H_{\text{bond breaking}}$
$= (+436) + (+242) = +678$ kJ

$\Delta H_{\text{bond making}}$
$= 2 \times (-431) = -862$ kJ

Overall ΔH for the reaction
$= (+678) + (-862) = -184$ kJ

Questions

1. In which equations are the signs of enthalpy change (ΔH) correctly shown?

	Equation	ΔH
a)	$CO_2(s) \longrightarrow CO_2(g)$	positive
b)	$CH_4(g) \longrightarrow C(g) + 4H(g)$	positive
c)	$6CO_2(g) + 6H_2O(l) \longrightarrow C_6H_{12}O_6(aq) + O_2(g)$	negative
d)	$2AgBr(s) \longrightarrow 2Ag(s) + Br_2(g)$	negative
e)	$CuSO_4(s) + 5H_2O(l) \longrightarrow CuSO_4.5H_2O(s)$	negative

2. The reaction between nitrogen and oxygen to form nitric oxide (NO) is an endothermic reaction.

$$\tfrac{1}{2}N_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow NO(g) \qquad \Delta H = +90 \text{ kJ}$$

a) Draw the energy level diagram to represent the energy change in this reaction.
b) What is the ΔH value for the reaction:

$$2NO(g) \longrightarrow N_2(g) + O_2(g) \qquad \Delta H = ?$$

17.3 | Activation Energy and Energy Profile Diagrams

Activation Energy

All reactions need energy in order to get started. For example, a mixture of hydrogen and oxygen will not explode unless it is ignited. Many reactions that do not occur at room temperature will proceed rapidly when the reactants are heated. Heat provides the energy to change less energetic particles into more energetic particles and thus to start a reaction.

The minimum energy that reacting particles must possess in order for a chemical reaction to occur is called the **activation energy, E_a**. A reaction will only occur when reacting particles possess activation energy.

A reaction involves reacting particles colliding to form product particles. Reacting particles that have less energy than the activation energy cannot break their bonds on collision with other reacting particles. Therefore, a reaction to form product particles does not occur.

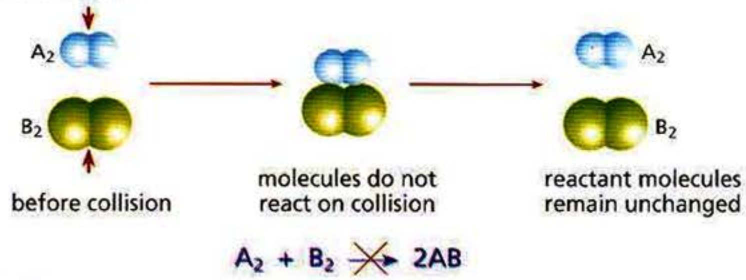


Fig. 17.10 No reaction occurs because the reactant molecules possess energy less than the activation energy.

More energetic particles are able to break their bonds on collision and form new bonds (Fig. 17.11). This is because these particles possess activation energy, which is the energy necessary for initiating a reaction.

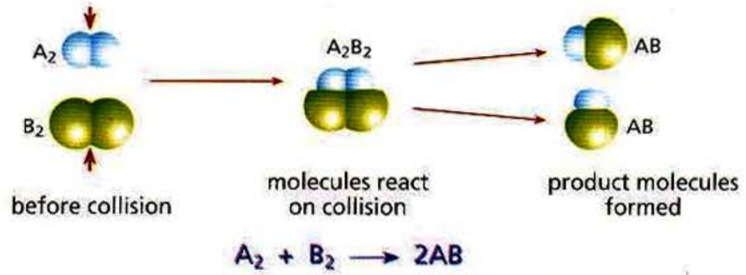


Fig. 17.11 A chemical reaction occurs because the reactant molecules possess activation energy.

Energy Profile Diagrams for Exothermic and Endothermic Reactions

To show the activation energy of a reaction, diagrams called **energy profile diagrams** are used.

An energy profile diagram is a way of representing the energy changes that occur during a chemical reaction. The energy difference between the products and reactants represents the enthalpy change of the reaction.

Fig. 17.12 shows the energy profile diagram of an exothermic reaction while Fig. 17.13 shows that of an endothermic reaction.

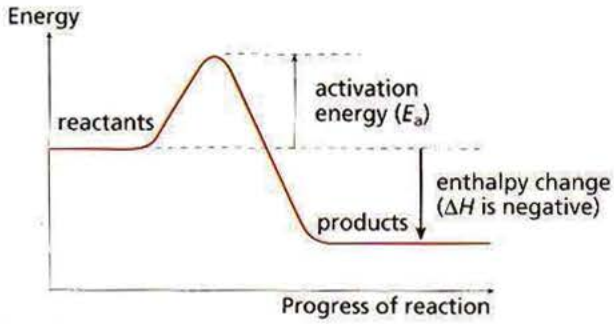


Fig. 17.12 Energy profile diagram of an exothermic reaction

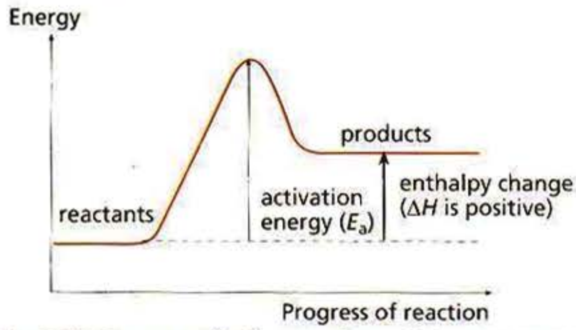


Fig. 17.13 Energy profile diagram of an endothermic reaction

17.4 | Combustion of Fuels

Every day, we use an enormous amount of energy in our homes, schools and factories for different purposes. We use energy to light up our homes and schools. We use energy to power our cars. We need energy in factories to turn raw materials into manufactured products. Most of the energy required is produced from the combustion of chemicals called **fuels**.

Key ideas

1. The minimum energy that particles must possess in order for a chemical reaction to occur is called the activation energy.
2. An energy profile diagram shows the activation energy required and the enthalpy change for a reaction.

What are fuels?

Fuels are *substances that can burn easily in air to give out energy*. Wood, coal, petroleum, hydrogen and natural gas are examples of fuels. The most commonly used fuels are **fossil fuels** such as coal, petroleum and natural gas. They were formed from decayed plants and animals that lived millions of years ago. They are carbon compounds.

What happens when fuels burn?

Coal is mainly carbon. It burns in an excess of air to give carbon dioxide and energy.

carbon + oxygen $\longrightarrow$ carbon dioxide + heat energy

$C(s) + O_2(g) \longrightarrow CO_2(g)$ + heat energy

Most fuels contain carbon and hydrogen. When these fuels burn, carbon dioxide, water and heat energy are produced. The equation for the combustion of natural gas, which contains mainly methane (CH_4), can be written as:

methane + oxygen $\longrightarrow$ carbon dioxide + water + heat energy

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$ + heat energy

This apparatus is called a bomb calorimeter. It is used to determine the heat produced when a substance burns.

If a limited supply of air is used, carbon particles (in the form of soot) and a poisonous gas, carbon monoxide, are produced. This is called **incomplete combustion**.

The equation for this reaction is:

methane + oxygen → carbon monoxide + water vapour + carbon + heat energy

$$4CH_4(g) + 5O_2(g) \longrightarrow 2CO(g) + 8H_2O(g) + 2C(s) + \text{heat energy}$$

All combustion reactions give out heat energy. The combustion of fuels is thus an exothermic process.

What is a good fuel?

Fuels are burnt to provide us with energy. But not every chemical which burns is a good fuel. The following chart shows us the criteria for selecting a good fuel.

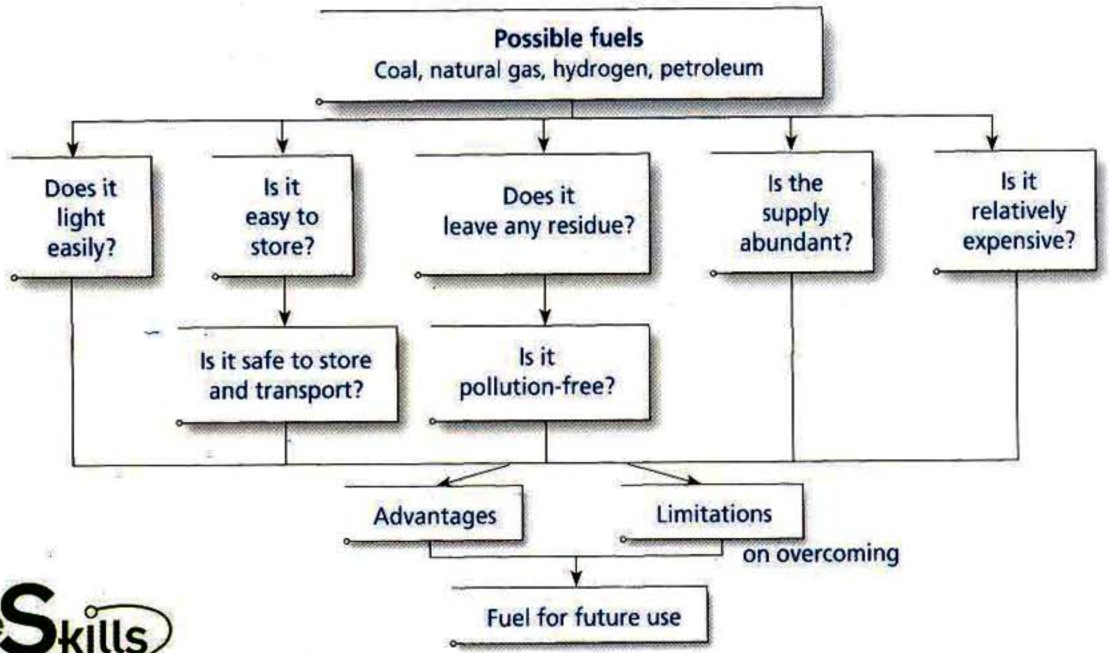


Fig. 17.14 Flow chart for evaluating a fuel

Thus, a good fuel:
- must release a lot of heat energy when it is burnt.
- must be safe to use and convenient to store.
- must be reasonably cheap and readily available.
- must not produce smoke, unpleasant gases or poisonous gases.

Science Skills

Before something can start burning, three things are needed: fuel, heat and oxygen. By removing one of them, the fire can be put out. Using this concept, explain why water is often used to put out fires.

Fire triangle

Fuel Cells

Combustion reactions are only one means of extracting useful energy from fuels. For more than a century, scientists have explored the possibility of converting the chemical energy of fuels directly to electricity. *A chemical cell in which reactants (usually a fuel and oxygen) are continuously supplied to produce electricity directly* is called a **fuel cell**.

The best known example is the hydrogen-oxygen fuel cell which is used as a source of electrical power in space vehicles (Fig. 17.15).

A fuel cell is a chemical cell. At the positive electrode (cathode), oxygen is reduced to form hydroxide ions. At the negative electrode (anode), hydrogen is oxidised to form water.

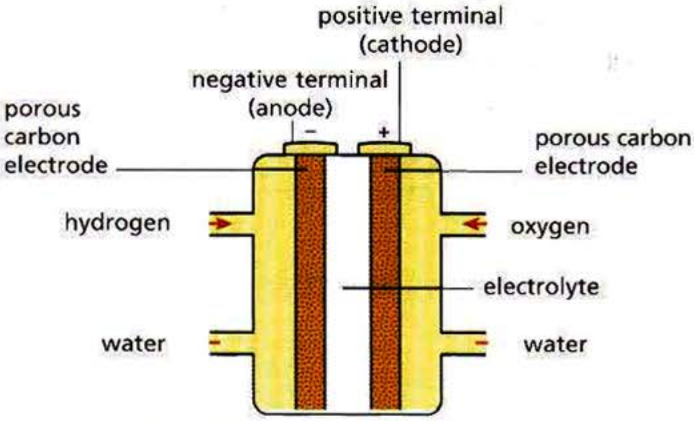


Fig. 17.15 The hydrogen-oxygen fuel cell

At the cathode: $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$

At the anode: $2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + 4e^-$

Overall reaction: $O_2(g) + 2H_2(g) \longrightarrow 2H_2O(l)$

The overall reaction is simply the conversion of hydrogen and oxygen to water. The reaction is the same as that for the combustion of hydrogen.

A fuel cell differs from an ordinary chemical cell (or battery) in one important aspect, i.e. the reactants are not contained within the cell but instead are continuously supplied from an external reservoir.

Is hydrogen the fuel for the future?

The world's reserves of fossil fuels are rapidly declining because the demand for them has increased exponentially. Hence, hydrogen may emerge as a potential fuel for the future. It is an ideal fuel in many ways. It is a renewable energy resource. It is also pollution-free because only water is produced when hydrogen reacts with oxygen. The basic problems are in finding a cheap source of hydrogen and an effective means of storing the gas.

A car designed to run on hydrogen-oxygen fuel cells

There is plenty of hydrogen on this planet, but it is mostly combined with oxygen in seawater. One possible way of obtaining hydrogen is by the electrolysis of water. Unfortunately, this is a very expensive method. Another alternative is to obtain hydrogen from hydrocarbons. In chapter 23, we shall study a process called cracking. Cracking breaks up hydrocarbons that produces hydrogen and other substances.

Link

What happens when acidified water is electrolysed? Recall what you have learnt in chapter 15.

Key ideas

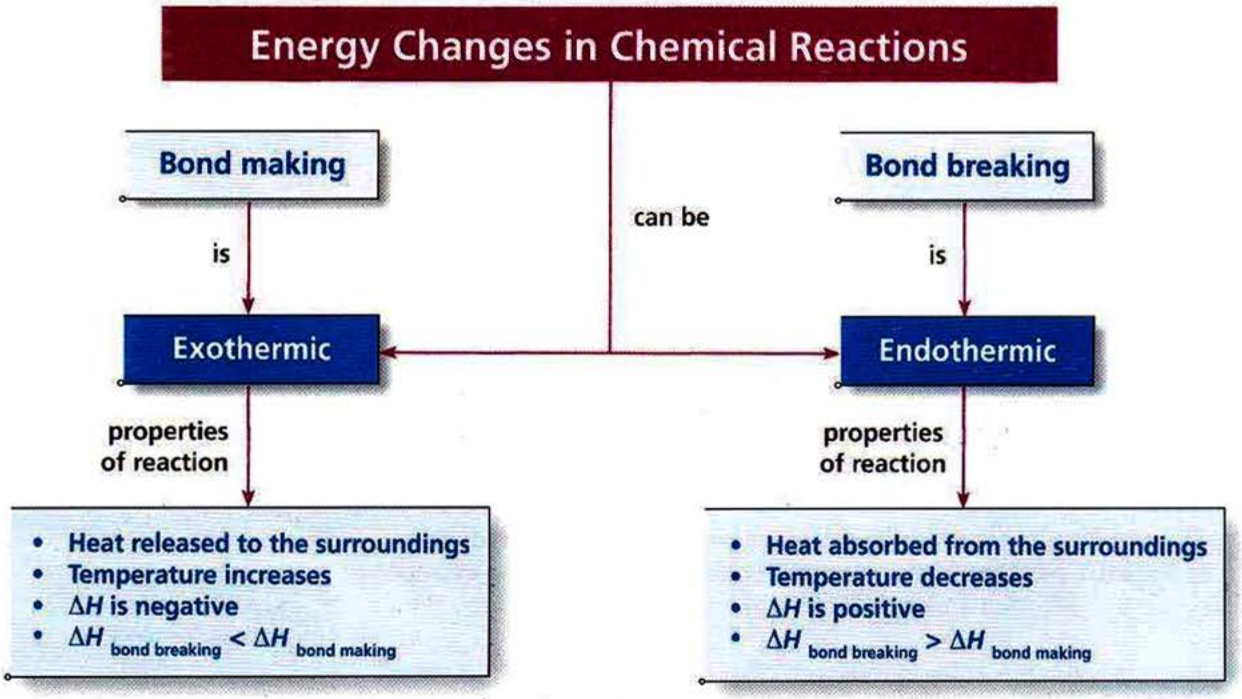
1. Fuels give out energy when burnt. The combustion of fuels is an exothermic reaction.
2. A fuel cell is a chemical cell in which reactants (usually a fuel and oxygen) are continously supplied to produce electricity directly.
3. Hydrogen, obtained from water or hydrocarbons, is a potential fuel for the future.

Test Yourself 17.3

Question

Give **two** similarities and **one** difference between combustion and rusting.

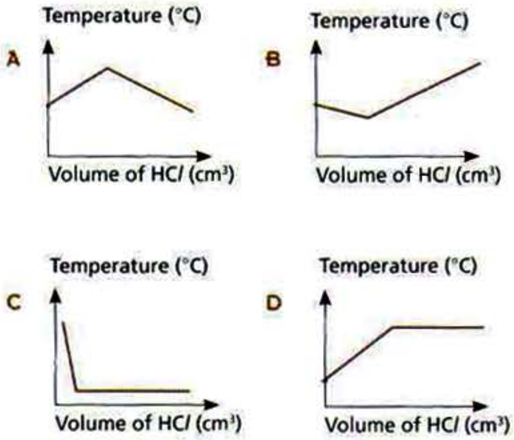
Concept Map



Exercise 17

Foundation

1. Which graph shows how the temperature changes as hydrochloric acid is added slowly to aqueous sodium hydroxide until the acid is in excess?



2. Which of the following processes are endothermic?

   1 Oxidation of carbon to carbon dioxide.
   2 Dissolving of anhydrous copper(II) sulphate in water.
   3 Adding of water to ammonium nitrate.
   4 Formation of carbohydrate and oxygen from carbon dioxide and water.

   A 1 and 2 only. B 3 and 4 only.
   C 1, 2 and 3 only. D 2, 3 and 4 only.

3. Which of the following processes is/are exothermic?
   Process 1: $I_2(g) \longrightarrow 2I(g)$
   Process 2: $CuCO_3(s) \longrightarrow CuO(s) + CO_2(g)$
   Process 3: $C(s) + O_2(g) \longrightarrow CO_2(g)$

   A 1 only. B 1 and 2 only.
   C 3 only. D 2 and 3 only.

4. Which of the following are endothermic processes?

   Process 1: $H_2(g) \longrightarrow 2H(g)$
   Process 2: $HCl(g) \longrightarrow H(g) + Cl(g)$
   Process 3: $H_2O(l) \longrightarrow H_2O(g)$
   Process 4: $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

   A 1 and 3 only.
   B 2 and 4 only.
   C 1, 2 and 3 only.
   D 1, 2, 3 and 4.

5. When nitrogen reacts with hydrogen to form ammonia, heat is given off.

   $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92$ kJ

   What is the value of ΔH for the following reaction?

   $NH_3(g) \rightleftharpoons \frac{1}{2}N_2(g) + \frac{3}{2}H_2(g)$

   A −92 kJ B −46 kJ
   C +46 kJ D +92 kJ

6. The formation of hydrogen chloride from hydrogen and chlorine is a strongly exothermic reaction. From this, it can be deduced that ________.

   A the temperature falls during the reaction
   B heat is absorbed in the reaction
   C hydrogen chloride gas burns readily in air
   D hydrogen chloride is not easily decomposed by heat

7. When hydrogen and iodine combine to form hydrogen iodide, heat is absorbed. From this, we can deduce that ________.

   A hydrogen and iodine react more rapidly at lower temperatures
   B the energy involved in making the bonds is less than that involved in breaking the bonds
   C more bonds are broken than formed
   D the formation of H–I bonds uses up energy

8. Bond energy is the energy required to break the covalent bond between two atoms.

   Based on the data given, which of the following statements is correct?

   A It is easier to break a C = C bond than a C – C bond.
   B The C – C bond is a stronger bond than the C = C bond.
   C The C = C bond is a stronger bond than the C – C bond.
   D The C = C bond is twice as strong as the C – C bond.

Bond	Bond energy (kJ/mol)
C – C	348
C = C	612

9. In the hydrogen-oxygen fuel cell, ________.
   A hydrogen reacts with oxygen to generate electricity
   B hydrogen ions react with hydroxide ions to
   D generate electricity
   C electricity is used to generate hydrogen and oxygen
   D electricity is used to provide heat energy

10. Anhydrous sodium carbonate and potassium nitrate were dissolved separately in water at 25 °C. The temperature of each solution was measured as soon as all the solid had dissolved.

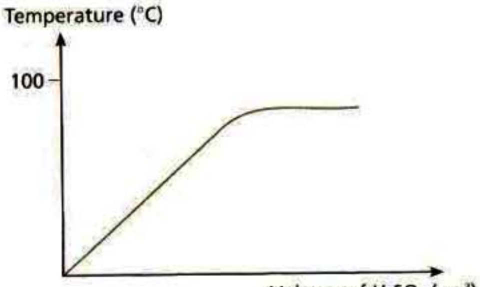
Compound		Anhydrous sodium carbonate	Potassium nitrate
Chemical formula			
Temperature (°C)	**Water**	25	25
	Solution	34	19
Temperature change (°C)			

a) Copy and complete the table to show
   i) the chemical formula of the compounds.
   ii) the temperature changes after dissolving the compounds.

b) i) State the type of energy change that takes place when anhydrous sodium carbonate dissolves in water.
   ii) Give a reason for your answer.

Challenge

1. Concentrated sulphuric acid was added slowly to water and the mixture constantly stirred. The temperature of the mixture was recorded at regular intervals. The results of the experiment are shown in the graph.



a) What can you infer from the graph?

b) Why should water never be added to concentrated sulphuric acid?

2. Hydrogen reacts with chlorine as shown by the equation below.

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g) \quad \Delta H = -184\ kJ$$

Calculate the energy change when 14.2 g of chlorine reacts completely with hydrogen.

3. a) When 50.0 cm³ of 1.0 mol/dm³ nitric acid is added to 50.0 cm³ of 1.0 mol/dm³ sodium hydroxide solution, x kJ of heat is evolved. Explain why heat is evolved.

   b) 50 cm³ of 1.0 mol/dm³ of nitric acid is added to 25 cm³ of 1.0 mol/dm³ potassium hydroxide solution.
      i) Calculate the number of moles of water molecules formed.
      ii) Based on your answer in (a), predict the amount of heat evolved (in terms of x) in this reaction.

4. When hydrogen burns in air, water is formed. The equation for the reaction is:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \qquad \Delta H = -484\ kJ$$

   a) Which has the greater amount of energy, water or the mixture of hydrogen and oxygen? Give a reason for your answer.

   b) The energy needed for forming the bonds of two water molecules is −1852 kJ. What is the energy needed for breaking the bonds of two hydrogen molecules and one oxygen molecule?

   c) Draw a labelled energy profile diagram for the reaction between hydrogen and oxygen to form water.

5. a) Define the term 'fuel cell'.
   b) Give **two** reasons why hydrogen is considered as a potential fuel for use in the future.
   c) What is the overall reaction in the hydrogen-oxygen fuel cell?
   d) Sketch the energy profile diagram for the reaction in hydrogen-oxygen fuel cell.
   e) Explain why this reaction is exothermic in terms of bond breaking and bond forming.

Chemistry Today

Food provides energy for our bodies. Each type of food has a caloric value measured in kilocalories or kcal. For example, carbohydrates: 4.0 kcal (17 kJ) of energy per gram; proteins: 4.0 kcal (17 kJ) of energy per gram; fat: 9.0 kcal (38 kJ) of energy per gram.

The energy provided by the food we eat is either used up or stored in our bodies. Thus,
energy intake = energy used up + energy stored

Our bodies use energy for maintaining metabolism (i.e. the chemical processes occurring in our bodies to keep us alive) and physical activities. The energy used up is also called the body's energy output. If the body's energy intake is less than its energy output, weight is lost.

CRITICAL THINKING

A particular milk drink has the following nutritional information:

100 ml of this drink contains	
carbohydrate	: 16 g
protein	: 4 g
fat	: 8 g

What is the percentage of total calories obtained from fat?

Chapter 18

Speed of Reaction

Zinc roof

This zinc roof has been worn away by acid rain over many years. Yet, in the laboratory, we can see that zinc dissolves in an acid within seconds! Both are chemical reactions but they take place at very different speeds. All around us, we see different chemical reactions occurring every day. Some occur slowly while others are complete in less than a second. Many reactions take place at speeds we can measure in the school laboratory.

Chemists also need to measure and know how fast a chemical reaction proceeds. This helps them to decide on the fastest and cheapest way to produce chemicals for use in the laboratory or in industry.

Chapter Outline

18.1 | Different Speeds of Reaction

Different chemical reactions take place at different speeds. When silver nitrate solution is added to sodium chloride solution, a white precipitate is produced immediately. This is a very fast reaction. However, the corrosion of silverware is a much slower reaction. Table 18.1 gives some other examples of reactions that proceed at different speeds.

Very fast	Moderately fast	Slow
• explosion of a petrol-air mixture • precipitation reactions	• reaction of metals or carbonates with dilute acids	• rusting of iron in air • reaction of magnesium with cold water • fermentation (conversion of fruit juice into alcohol)

Table 18.1 Different reactions proceed at different speeds

How do we calculate the speed of a reaction?

During a chemical reaction, the reactants get used up as the products are formed. We can measure the **speed of a reaction** by measuring the *amount of a reactant used up per unit time*, i.e.

$$\text{Speed of reaction} = \frac{\text{amount of reactant used up}}{\text{time taken}}$$

The speed of a reaction can also be measured in terms of the *amount of a product obtained per unit time*, i.e.

$$\text{Speed of reaction} = \frac{\text{amount of product obtained}}{\text{time taken}}$$

For a chemical reaction that produces a gas, the speed of reaction can be found by measuring the *volume of gas produced per unit time*, i.e.

$$\text{Speed of reaction} = \frac{\text{volume of gas produced}}{\text{time taken}}$$

Key ideas

1. Different chemical reactions take place at different speeds.
2. The speed of reaction

$$= \frac{\text{change in amount of reactant or product}}{\text{time taken}}$$

OR (in some cases)

$$= \frac{\text{change in volume of gas}}{\text{time taken}}$$

TidBit

Why are the pages of old books fragile? Aluminium sulphate is often used to treat paper. Over time, the aluminium ions (Al^{3+}) in paper react with moisture in the air to produce hydrogen ions (H^+). Hydrogen ions break down the cellulose in paper and make paper brittle. Suggest methods that can be used to slow down this chemical reaction.

18.2 | Measuring Speed of Reaction in the Laboratory

The speed of a reaction can be found by measuring these quantities at regular time intervals:

- The volume of a gas produced by the reaction
- The mass of the reactant that remains

Measuring Speed of Reaction from Changes in Volume

Let us look at an example. The reaction between a reactive metal and a dilute acid is considered fast. For example, magnesium reacts with dilute hydrochloric acid according to the equation:

magnesium + dilute hydrochloric acid ⟶ magnesium chloride + hydrogen

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

As the reaction proceeds, the total volume of hydrogen gas produced increases. The speed of the reaction can therefore be determined by collecting and measuring the volume of hydrogen produced at regular time intervals (Fig. 18.1).

Experiment 1

To study the speed of reaction between dilute hydrochloric acid and magnesium

Procedure

1. The apparatus is set up as shown in Fig. 18.1. The layer of oxide on the magnesium ribbon is removed using a piece of sandpaper. This ensures that magnesium reacts with the dilute hydrochloric acid. The magnesium ribbon is then put in a small test tube.
2. The conical flask is shaken to mix the magnesium ribbon and acid. The stopwatch is started at the same time.
3. The volume of hydrogen collected in the gas syringe is recorded every half-minute.

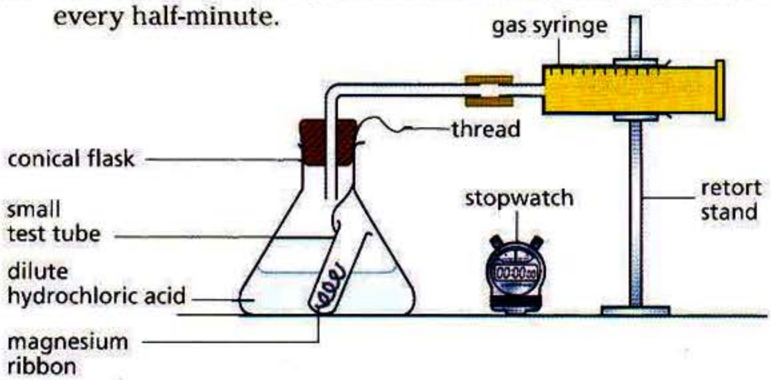


Fig. 18.1 An experiment to study the speed of reaction by measuring the volume of gas evolved

From the results of this experiment, a graph of the volume of hydrogen produced is plotted against time. The graph would look like the one shown in Fig. 18.2.

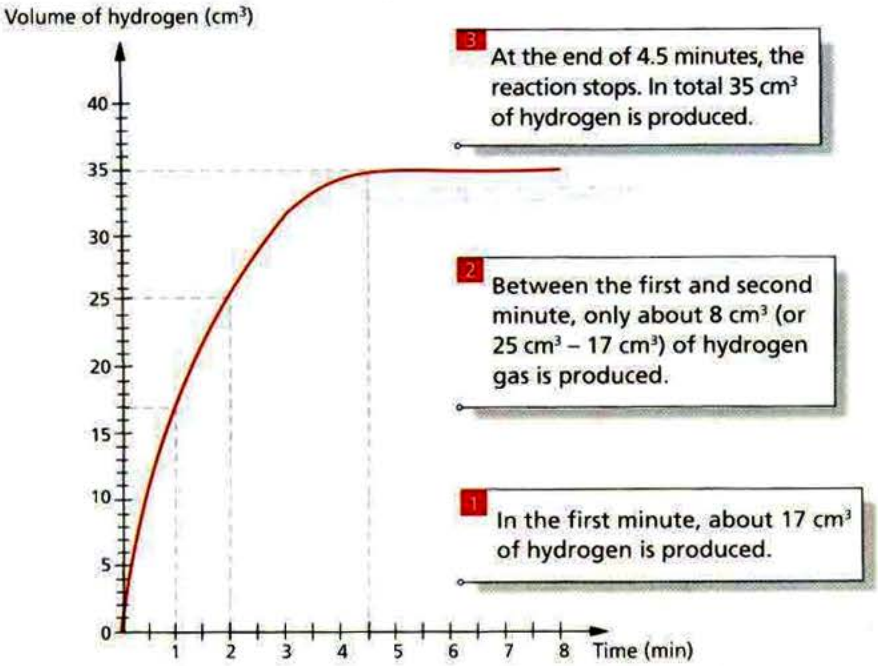


Fig. 18.2 A graph showing the volume of hydrogen produced at different time intervals

Science Skills

Does the speed of reaction increase, decrease or remain constant as the reaction proceeds?

How can we find the speed of reaction at a particular time from a graph?

Generally, for any reaction, if the change in the mass or volume of the reactants (or products) is plotted against time, the speed of the reaction is indicated by the gradient of the graph. Thus, in Fig. 18.3, the gradient at t_1 is the speed of reaction at that time.

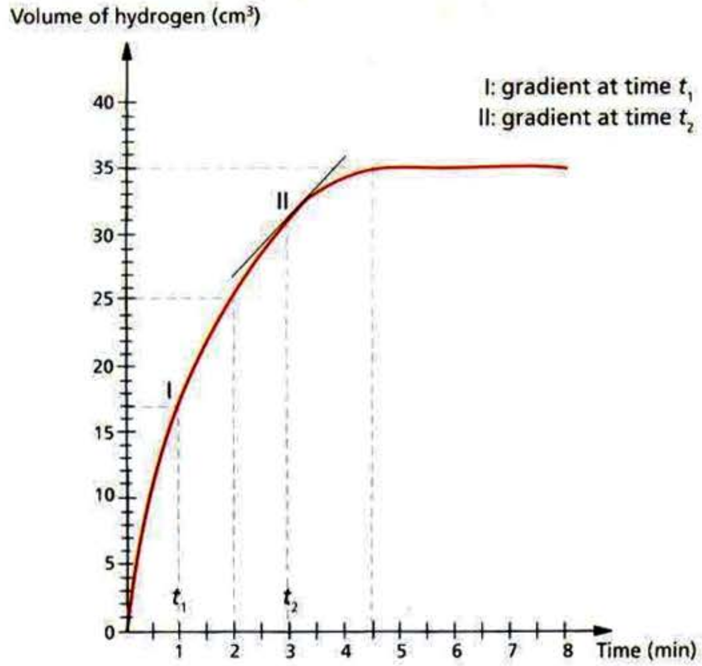
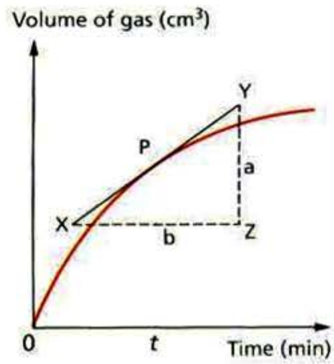


Fig. 18.3 The gradient of the graph at different times

Chem-Aid



The gradient at P (time(t)) can be determined as follows:

Step 1: Draw a tangent to the curve at P.

Step 2: Complete the triangle XYZ.

Step 3: Measure the distances, a (YZ) and b (XZ).

Step 4: Gradient at P = $\frac{a}{b}$

TidBit

This building was constructed from black marble and white marble. Calcium carbonate in marble buildings or structures react with the minute amounts of acid in rain to produce carbon dioxide. Over time, the marble wears away due to this reaction. Are there special methods that builders use to build marble structures in order to protect them against such damage?

How can we estimate the change of speed of reaction from a graph?

The shape of the graph tells us whether the speed of reaction changes or remains the same as time passes. The steeper the gradient, the faster the speed of reaction. For example, the graph in Fig. 18.3 shows that the gradient at t_2 (gradient II) is less steep than that at t_1 (gradient I). As the reaction proceeds, the curve becomes less steep. This means that the speed of reaction is decreasing.

Measuring Speed of Reaction from Changes in Mass

The speed of a reaction can also be found by measuring the changes in mass of a reaction mixture. This method works best for reactions which produce gases such as carbon dioxide. For example, the speed of reaction between calcium carbonate and hydrochloric acid can be studied this way. The apparatus used is shown in Fig. 18.4.

Experiment 2

To study the speed of reaction between calcium carbonate and dilute hydrochloric acid

Fig. 18.4 An experiment to study the speed of reaction by measuring the mass at different time intervals

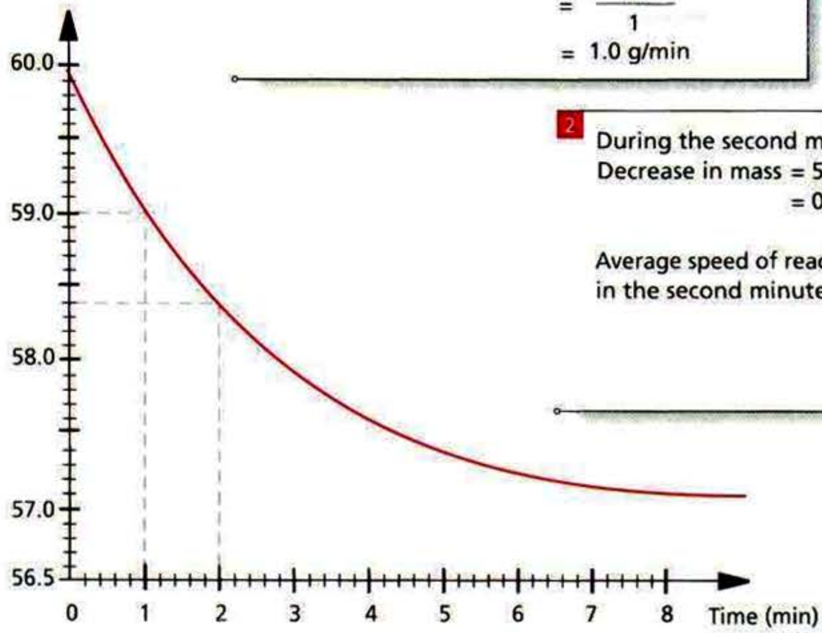
Procedure

1. The apparatus is set up as shown in Fig. 18.4. The cotton wool in the mouth of the conical flask is used to prevent acid spray i.e. to stop the acid from splashing out as the reaction takes place.
2. The mass of the system is recorded. This includes the mass of the marble chips (calcium carbonate), dilute hydrochloric acid, conical flask, small test tube, string and cotton wool.
3. The conical flask is shaken to mix the marble chips and acid. The stopwatch is immediately started.
4. The mass of the system is recorded at one-minute intervals.

A sample set of results from experiment 2 is used to plot a graph as shown in Fig. 18.5.

1 During the first minute,
Decrease in mass = 60.0 g − 59.0 g
= 1.0 g

$$\text{Average speed of reaction in the first minute} = \frac{\text{decrease in mass}}{\text{time taken}} = \frac{1.0}{1} = 1.0 \text{ g/min}$$



2 During the second minute,
Decrease in mass = 59.0 g − 58.4 g
= 0.6 g

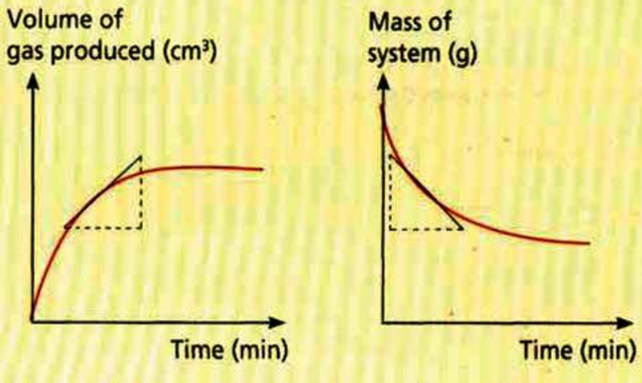
$$\text{Average speed of reaction in the second minute} = \frac{\text{decrease in mass}}{\text{time taken}} = \frac{0.6}{1} = 0.6 \text{ g/min}$$

3 The gradient decreases with time, i.e. the reaction slows down as it proceeds. The reaction stops after 7.5 minutes.

Fig. 18.5 A graph showing the mass of the system at different time intervals

Key ideas

1. The speed of reaction at any instant is indicated by the gradient of graphs shown:



2. The speed of a reaction decreases as the reaction proceeds.

Test Yourself 18.1

Worked Example

Excess zinc is added to 50.0 cm³ of 1.0 mol/dm³ sulphuric acid. The reaction is complete in 80 seconds.

a) Calculate the total volume of hydrogen gas released.
b) Sketch a graph to show the volume of hydrogen gas released against time.

(1 mol of gas occupies 24 dm³ at r.t.p.)

Thought Process

a) The reaction between zinc and sulphuric acid produces hydrogen gas.

Step 1: Calculate the number of moles of sulphuric acid used.
Step 2: Write the balanced chemical equation between sulphuric acid (H_2SO_4) and zinc (Zn).
Step 3: From steps 1 and 2, calculate the number of moles of hydrogen gas produced.
Step 4: From step 3, calculate the volume of hydrogen gas (H_2) produced.

b) Sketch the graph to show the total volume of hydrogen gas produced in 80 seconds. As with most reactions, the speed of reaction decreases as the reaction proceeds. This means the gradient of the graph is becoming less steep as time passes. Thus, the graph is curved as shown.

Answer

a) Amount of H_2SO_4 used = $\frac{1.0 \times 50}{1000}$

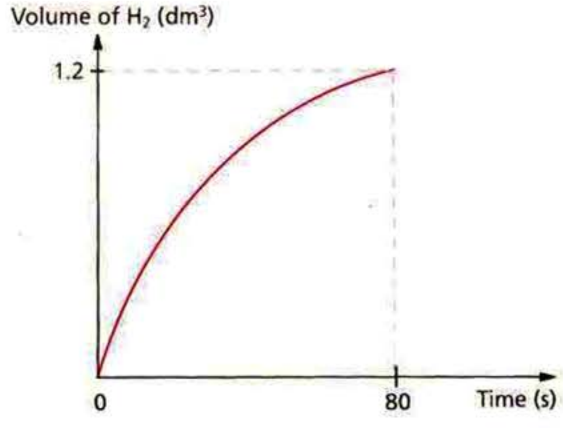
= 0.05 mol ---------------------- (step 1)

$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$ ------------- (step 2)

1 mol of H_2SO_4 produces 1 mol of H_2.
∴ 0.05 mol of H_2SO_4 produces 0.05 mol of H_2. ------------ (step 3)

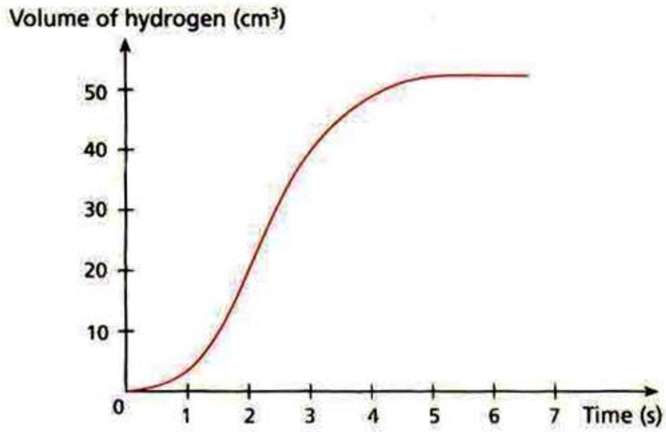
1 mol of H_2 occupies 24 dm³ at r.t.p.
∴ 0.05 mol of H_2 occupies $0.05 \times 24 = 1.2$ dm³ at r.t.p. ------ (step 4)

b)



Question

The graph shows the results of an experiment involving the reaction between magnesium ribbon and excess dilute hydrochloric acid.



a) Was the reaction slow or fast at first? Suggest one reason for this observation.
b) How much time did the magnesium ribbon take to react completely?
c) At which time was the reaction fastest?
d) Another similar reaction produced 12 cm^3 of hydrogen in the first second. Is this reaction faster than the previous reaction?
e) Can the change in the concentration of hydrochloric acid be used to measure the rate of this reaction? Explain why.

18.3 | Factors Affecting Speed of Reaction

Many factors affect the speed of a chemical reaction. These include

- the concentration of the reactants,
- the pressure of the reactants (if the reactants are gaseous),
- the particle size or surface area of the reactants,
- the temperature at which the reaction is occurring.

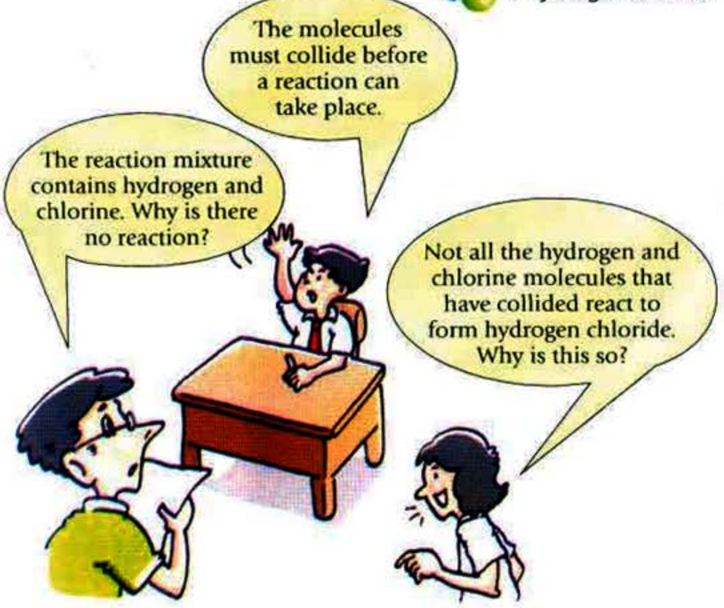
Why do these factors affect the speed of a chemical reaction? To answer this, we must study what happens to the particles (atoms, molecules, ions, etc.) of the reacting substances.

Consider the reaction between hydrogen (H_2) and chlorine (Cl_2). Molecules of hydrogen and chlorine react to form molecules of hydrogen chloride only when exposed to sunlight.

hydrogen + chlorine $\xrightarrow{\text{sunlight}}$ hydrogen chloride

$H_2(g) + Cl_2(g) \xrightarrow{\text{sunlight}} 2HCl(g)$

Fig. 18.6 A mixture of H_2 and Cl_2 gases



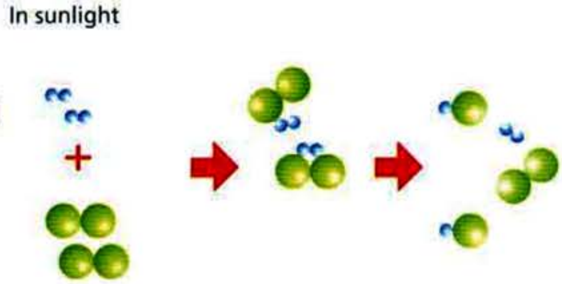


Fig. 18.7 Only some of the molecules that collide react to form HCl.

For a reaction to occur between two particles:

1. the reacting particles must collide with each other,
2. they must collide with a certain *minimum amount of energy* known as the **activation energy**.

Link

Particles of matter are in constant random motion. Particles with more energy move faster. Recall what you have learnt in chapter 1.

In this way, collisions between reacting particles result in the formation of product particles. These collisions are known as **effective collisions**. Thus, in the reaction between hydrogen and chlorine, only fast-moving molecules with energies equal to or greater than the activation energy will react on collision to form hydrogen chloride.

The collision of reacting particles can be used to explain why the speed of a reaction varies with concentration, pressure, particle size and temperature.

In general, *when any factor increases the rate of effective collisions between reacting particles, it will also increase the speed of reaction.*

Effect of Concentration on Speed of Reaction

When a reaction involves solutions, we can alter the speed of reaction by changing the concentration of the reactant. To study the effect of concentration on the speed of reaction, let us look at the reaction between magnesium ribbon and dilute hydrochloric acid.

Experiment 3

To study the effect of concentration on the speed of the reaction between magnesium and hydrochloric acid

Procedure

1. The apparatus is set up as shown in Fig. 18.8.

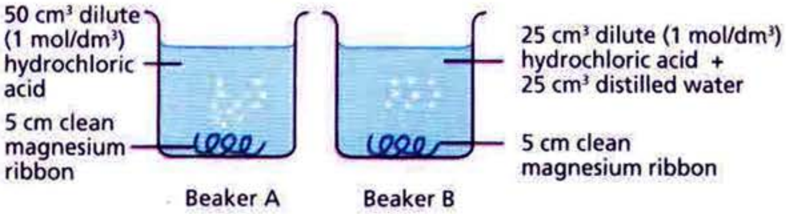


Fig. 18.8 An experiment to study the effect of concentration on the speed of reaction

2. The time taken for each piece of magnesium ribbon to dissolve completely is recorded.

The results of this experiment are shown in Table 18.2.

Beaker	A	B
Time taken for magnesium to dissolve (s)	39	78

Table 18.2 Results of experiment 3

In the above experiment, the acid in beaker A is twice as concentrated as the acid in beaker B. The time taken for magnesium to react completely in beaker A is shorter. This means that magnesium reacts faster in the more concentrated acid. Thus, we can conclude that a reaction proceeds faster when the concentration of a reactant is increased.

We can study the reaction between magnesium and hydrochloric acid in more detail using the apparatus shown in Fig. 18.9.

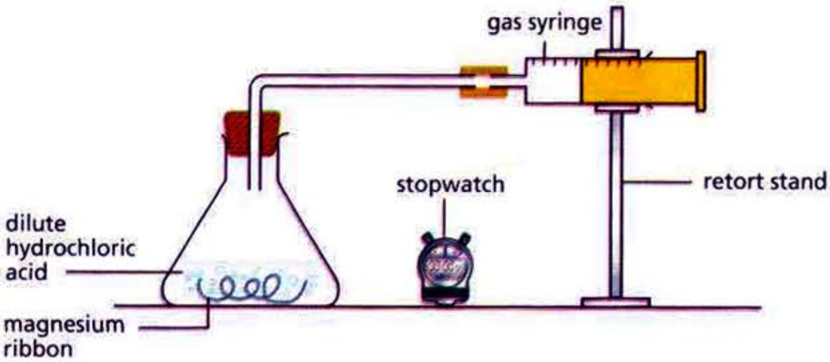


Fig. 18.9 An experiment to study the effect of concentration on the speed of reaction by measuring the volume of gas evolved

Science Skills

1. Compare the hydrochloric acid in beakers A and B.
   a) Which beaker contains a more concentrated acid?
   b) How did you come to this conclusion?
2. Which reaction proceeds faster, the reaction in beaker A or beaker B? Explain your answer.

Chem-Aid

1. There are many factors that affect the speed of reaction. In this experiment, the following factors are kept constant:
   - Size of magnesium ribbon
   - Temperature at which the experiment was carried out
   - Volume of solution in each beaker
2. The only variable that is changed is the concentration of hydrochloric acid.

The experimental details for the reaction between magnesium and hydrochloric acid are given in Table 18.3.

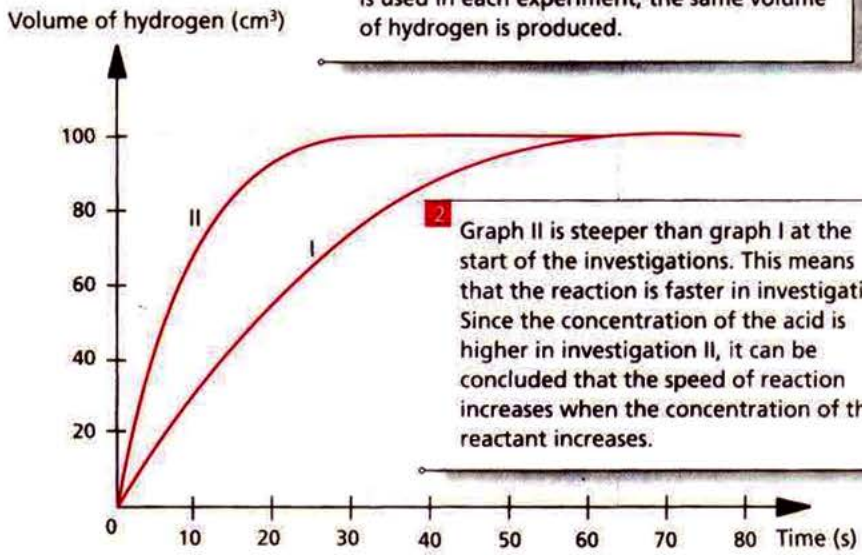
Experiment	I	II
Volume of hydrochloric acid (cm^3)	50.0	50.0
Concentration of hydrochloric acid (mol/dm^3)	1.0	2.0
Mass of magnesium (g)	0.1	0.1

Table 18.3 Experimental conditions to investigate the effect of concentration on speed of reaction

For both investigations, the volume of hydrogen produced is recorded at regular time intervals.

The results of the two investigations are plotted on the same axes, with the syringe readings on the y-axis and the time taken on the x-axis (Fig. 18.10).

1 In both investigations, hydrochloric acid is in excess and magnesium is the limiting reactant. Since the same mass of magnesium is used in each experiment, the same volume of hydrogen is produced.



2 Graph II is steeper than graph I at the start of the investigations. This means that the reaction is faster in investigation II. Since the concentration of the acid is higher in investigation II, it can be concluded that the speed of reaction increases when the concentration of the reactant increases.

3 In both investigations, the reaction slows down and the graph becomes less steep over time. This is because the concentration of hydrochloric acid and the amount of magnesium decrease as the reaction proceeds.

Fig. 18.10 **Graphs tc show the effect of concentration on the speed of a reaction**

Why does the speed of reaction increase with concentration?
As a solution containing a reactant becomes more concentrated, there are more particles of the reactant occupying a given volume (Fig. 18.11). This means that the reacting particles can collide with one another more frequently. The more collisions there are, the more likely that particles collide with enough energy to react. Hence, the faster the reaction.

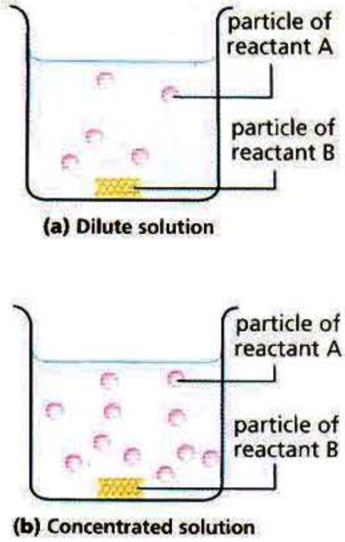


Fig. 18.11 A concentrated solution contains more particles of the liquid reactant A

Effect of Pressure on Speed of Reaction

Pressure has very little effect on reactions involving only solids or liquids. However, in chemical reactions that involve gases, increasing the pressure increases the speed of reaction.

What is the relationship between the pressure of a gas and its concentration?
In reactions involving gases, changing the pressure of a gas is similar to changing its concentration. At low pressures, the gaseous molecules are spread far apart (Fig. 18.12). At high pressures, the gaseous molecules are closer together (Fig. 18.13). This means that a gas at high pressure is more concentrated than a gas at low pressure. Collisions between the gaseous molecules become more frequent. Thus, increasing the pressure of the gas increases the speed of reaction.

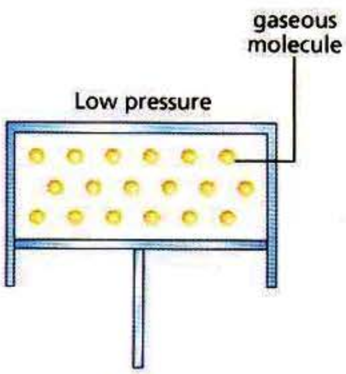


Fig. 18.12 At low pressure, particles are spread far apart.

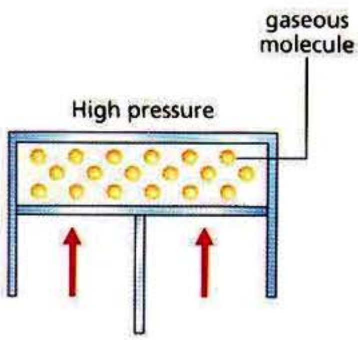


Fig. 18.13 At high pressure, the same number of particles are closer together.

In many industrial processes, such as the manufacture of ammonia from hydrogen and nitrogen (see chapter 19), high pressures (up to 250 atmospheric pressures) are used to speed up the chemical reactions.

Sometimes, the increase in pressure can make a reaction occur so fast that it becomes dangerous. This happens in coal mines where gases that can combust in air, such as methane, are present. As these gases accumulate, pressure is built up, causing explosions.

Key ideas

1. For reactions involving solutions, an increase in the concentration of a reactant increases the speed of reaction.
2. For reactions involving gases, an increase in the pressure of a gas increases the speed of reaction.
3. The effect of these factors on the speed of reaction can be explained by the collision of reacting particles.

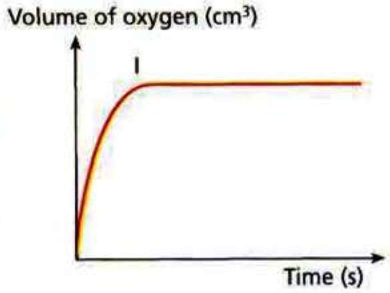
Test Yourself 18.2

Worked Example

The decomposition of hydrogen peroxide can be represented by the equation:

$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$

This reaction is carried out using 100 cm^3 of 1 mol/dm^3 hydrogen peroxide in an investigation (I). The results are plotted on the axes shown below.



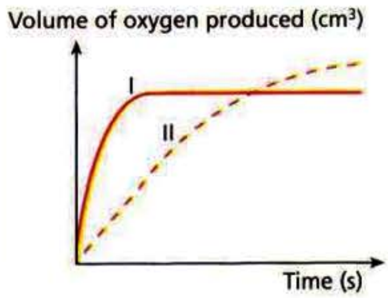
A second investigation (II) is carried out by adding a small amount of 0.1 mol/dm^3 hydrogen peroxide to 100 cm^3 of 1 mol/dm^3 hydrogen peroxide solution. Sketch the graph that you would expect to obtain for investigation II on the same axes.

Thought Process

The faster the speed of the reaction, the steeper the graph.

- The hydrogen peroxide in investigation II is more dilute than that in investigation I because a small amount of 0.1 mol/dm^3 hydrogen peroxide was added.
- The speed of reaction for investigation II is therefore slower and the graph is less steep.
- A larger volume of oxygen is produced in investigation II because more hydrogen peroxide was present. Thus, the graph would level off (to a horizontal line) above graph I.

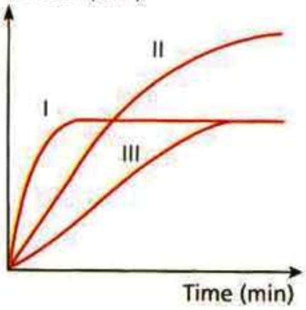
Answer



Questions

1. An iodine stain on a piece of cloth can be removed with sodium thiosulphate solution. Is it easier to remove the stain with a solution of concentration 1.0 mol/dm³ or 0.5 mol/dm³? Why?

2. Three investigations were carried out for the reaction between calcium carbonate powder and excess hydrochloric acid of different concentrations. The results of the investigations are shown below.



   a) Which investigation used hydrochloric acid of the lowest concentration?
   b) What do you notice about the speed of reaction for all three investigations as the reaction proceeds? Explain your observations.
   c) What can you infer from graph III?

Effect of Particle Size (or Surface Area) on Speed of Reaction

Many chemical reactions involve solids. Solids come in different sizes — we say they have different particle sizes. How is the speed of reaction affected by particle size? To find out, we can study the reaction between marble chips and dilute hydrochloric acid (experiment 4).

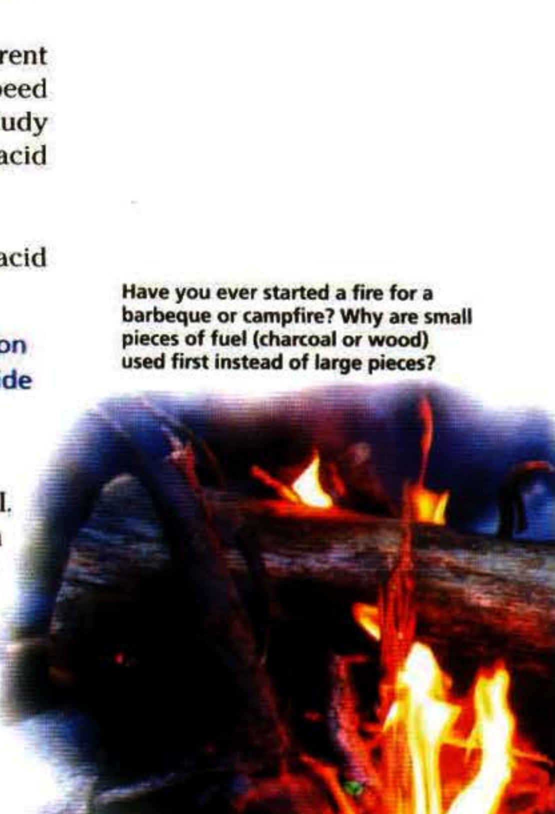
Marble chips, or calcium carbonate, react with hydrochloric acid according to the equation:

calcium carbonate + dilute hydrochloric acid ⟶ calcium chloride + water + carbon dioxide

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

The reaction is carried out twice (see page 332). In investigation I, large marble chips (large particle size) are used. But in investigation II, small marble chips (small particle size) are used. In each investigation, the same amounts of acid and calcium carbonate are used. The speed of reaction is determined by measuring the volume of carbon dioxide gas produced at regular time intervals.

Have you ever started a fire for a barbeque or campfire? Why are small pieces of fuel (charcoal or wood) used first instead of large pieces?



Chem-Aid

1. In this experiment, the following factors are kept constant:
   - Volume of hydrochloric acid
   - Concentration of hydrochloric acid
   - Temperature at which the experiment is carried out
   - Total mass of marble chips
2. The only variable that is changed is the size of the marble chips.

Experiment 4

To study the effect of particle size on the speed of reaction

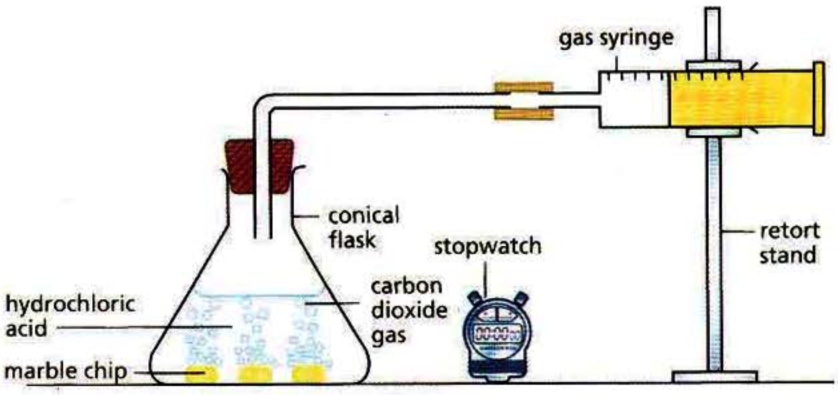


Fig. 18.14 An experiment to study the effect of particle size on the speed of reaction

Procedure

1. The experiment is set up as shown in Fig. 18.14. The volume of gas produced is recorded at one-minute intervals for investigation I.
2. The experiment is repeated for investigation II, with marble chips that have been crushed into much smaller pieces.

The results of the experiment are shown in Fig. 18.15.

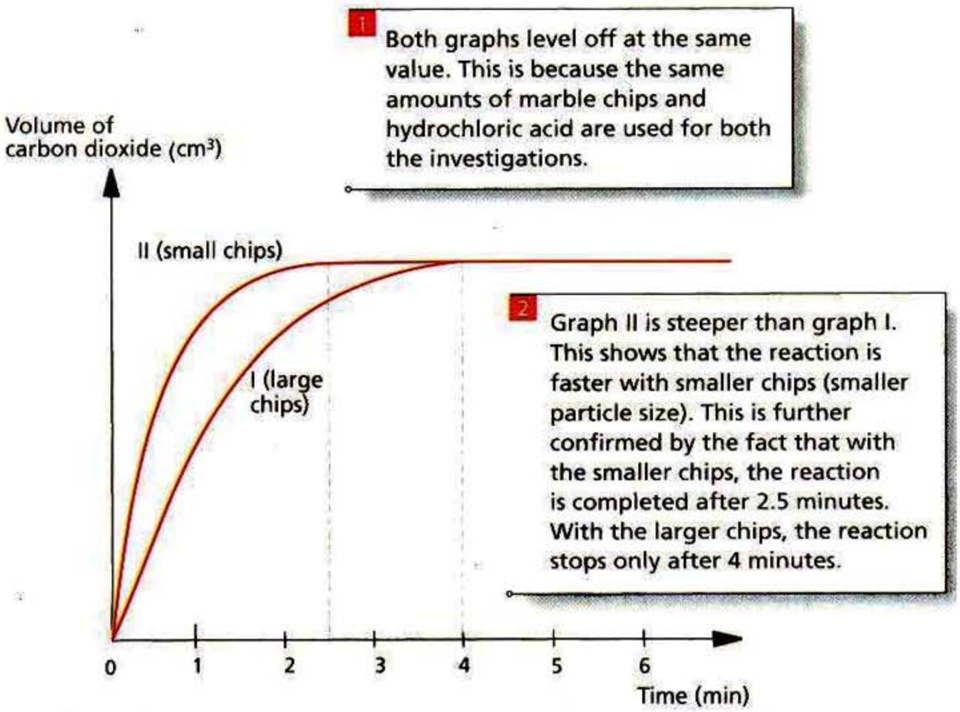


Fig. 18.15 Graphs to show the effect of particle size on the speed of reaction

The speed of the reaction between hydrochloric acid and calcium carbonate can also be measured using the data logger. This is done using the set-up shown in Fig. 18.16.

ScienceSkills

Both the effect of particle size and the effect of concentration on the speed of this reaction can be studied using data logging.

The rubber stopper may be blown out by the increase in pressure inside the conical flask. It is thus very important to wear safety goggles and use the quantities of chemicals as specified in the instructions for the experiment.

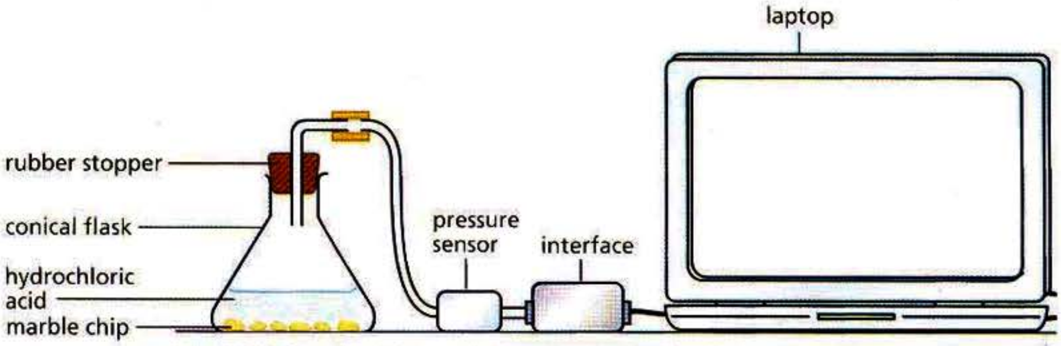
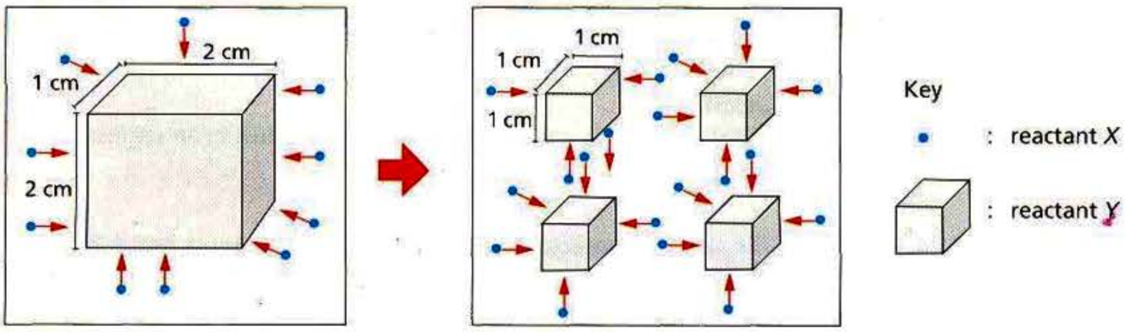


Fig. 18.16 Recording the speed of reaction by data logging

The increase in pressure is due to the carbon dioxide gas produced as the reaction proceeds. A pressure sensor measures the pressure increase in the reaction system.

Why does the speed of reaction increase with decreasing particle size (increasing surface area) of the solid reactant?

If a solid is cut into smaller pieces, its surface area gets larger. That is, extra surfaces are exposed for reactant particles to collide into (Fig. 18.17). This results in more reactions taking place in a shorter time.

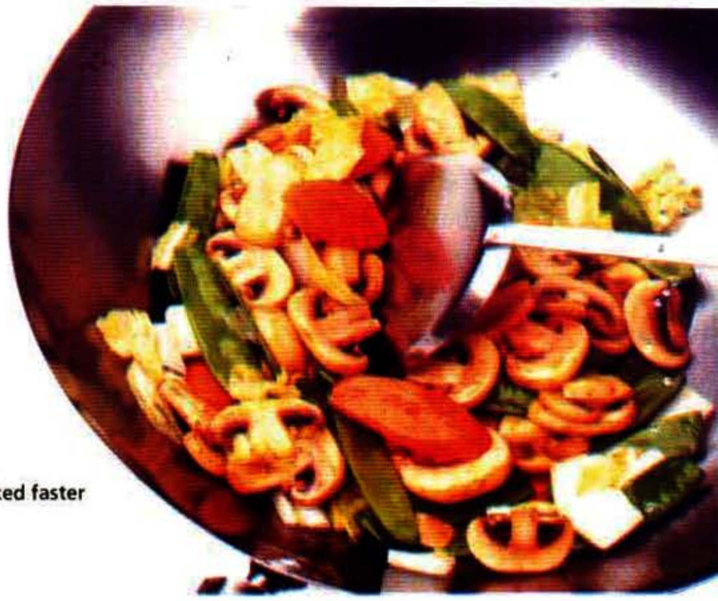


(a) Surface area = 16 cm²

(b) Surface area = 24 cm²

Fig. 18.17 Smaller particles of a solid reactant have a larger surface area available for reactions to occur on.

The small particles have a larger total surface area than a large particle of the same mass. Thus, *the smaller the particles, the greater the surface area, and the greater the speed of reaction.*



Small pieces of food can be cooked faster than larger pieces.

An example in daily life shows how particle size affects the speed of a reaction. Meat and vegetables contain chemicals such as proteins and carbohydrates. These chemicals are reactants in the chemical reactions that occur during the cooking process. Small pieces of food have a larger total surface area. Thus, they can be cooked faster and more thoroughly.

Key ideas

1. When a solid reacts with a liquid or a gas, the smaller the particle size (i.e. the larger the surface area) of the solid, the greater the speed of the reaction.
2. Reactions take place faster when the solid reactant is broken into smaller pieces.

Test Yourself 18.3

Questions

1. Coal mining can result in a build-up of coal dust in an enclosed area. Explain why this is dangerous.
2. Explain why antacid tablets for gastric patients work better if they are chewed before being swallowed.
3. Sketch two graphs to show the effect of particle size on the speed of reaction between dilute hydrochloric acid and copper(II) carbonate. Draw both graphs on the same axes (mass of reaction mixture against time).

Why does milk turn sour easily in warm surroundings?

Effect of Temperature on Speed of Reaction

Milk will turn sour very quickly if it is exposed to the air at room temperature, but it will keep fresh for several days in a refrigerator. The souring of milk is a decomposition reaction caused by bacteria. How is the rate of this reaction affected by temperature?

A chemical reaction can be made to proceed faster or slower by increasing or decreasing the temperature of the reactants. The most convenient reaction for studying this effect is the reaction between a dilute acid and sodium thiosulphate solution.

When dilute hydrochloric acid is added to sodium thiosulphate solution, a fine precipitate of sulphur slowly forms and the solution becomes cloudy. The reaction can be represented as:

sodium thiosulphate + hydrochloric acid ⟶ sodium chloride + water + sulphur + sulphur dioxide

$$Na_2S_2O_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$$

The speed of this reaction can be calculated as the rate of sulphur precipitation.

$$\text{Speed of reaction} = \frac{\text{amount of sulphur precipitated}}{\text{time taken}}$$

In the following experiment, the time is recorded for the same amount of sulphur to precipitate. The shorter the time taken for sulphur to precipitate, the higher the speed of reaction. This means the speed of reaction is *inversely proportional* to the time taken, i.e.

$$\text{Speed of reaction} \propto \frac{1}{\text{time taken}}$$

Experiment 5

To study the effect of temperature on the speed of reaction between sodium thiosulphate solution and dilute hydrochloric acid

Procedure

1. The apparatus is set up as shown in Fig. 18.18.

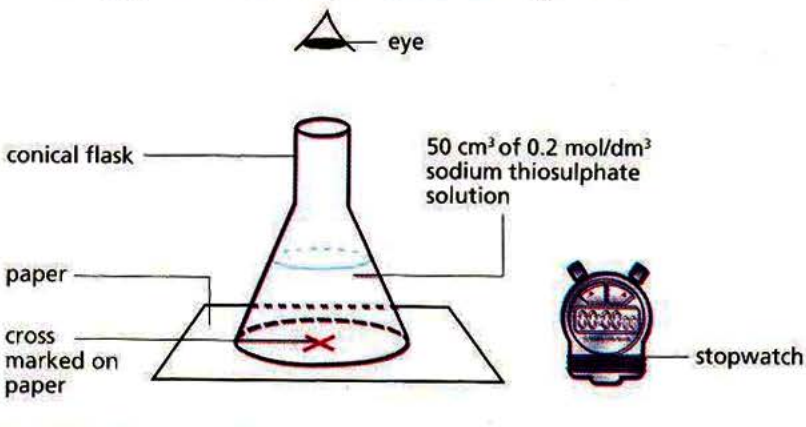


Fig. 18.18 An experiment to study the speed of reaction between hydrochloric acid and sodium thiosulphate

Chem-Aid

1. In this experiment, the following factors are kept constant:
   - Volume of hydrochloric acid and sodium thiosulphate solution
   - Concentration of hydrochloric acid and sodium thiosulphate solution
2. The only variable that is changed is the temperature at which the experiment is carried out.

2. 5 cm^3 of dilute hydrochloric acid is then quickly poured into the sodium thiosulphate solution and the stopwatch is started immediately.
3. The mixture is swirled once and the stopwatch is stopped at the moment the cross disappears from view. The time taken is recorded.
4. The experiment is repeated three times using fresh solutions of the same reactants, but with sodium thiosulphate solution heated to higher temperatures.

Fig. 18.19 shows the results obtained in experiment 5. In this experiment, the same amount of sulphur is precipitated at each temperature. This means that the less time it takes for the cross to disappear from view, the faster the reaction is.

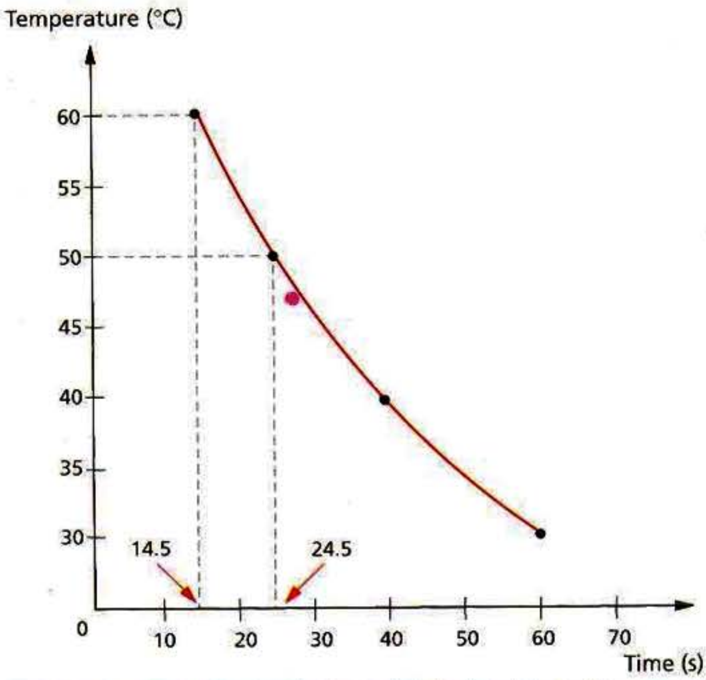


Fig. 18.19 A graph showing the time taken for the cross to disappear at different temperatures

Look carefully at the graph in Fig. 18.19.

1. The higher the temperature, the shorter the time taken for the cross to disappear from view. This means that the higher the temperature, the faster the speed of reaction.
2. Observe the gradient of the graph at various temperatures. The speed of reaction increases rapidly as the temperature increases.

To see how the speed of reaction varies with temperature more clearly, we can plot the results from experiment 5 using a different pair of axes. In this experiment, the speed of reaction is *inversely proportional* to the time taken for the cross to disappear from view, i.e.

$$\text{Speed of reaction} \propto \frac{1}{\text{time for cross to disappear from view}}$$

Thus, we can plot values of $\frac{1}{\text{time}}$ against temperature. It is the same as plotting the speed of reaction against temperature (Fig. 18.20).

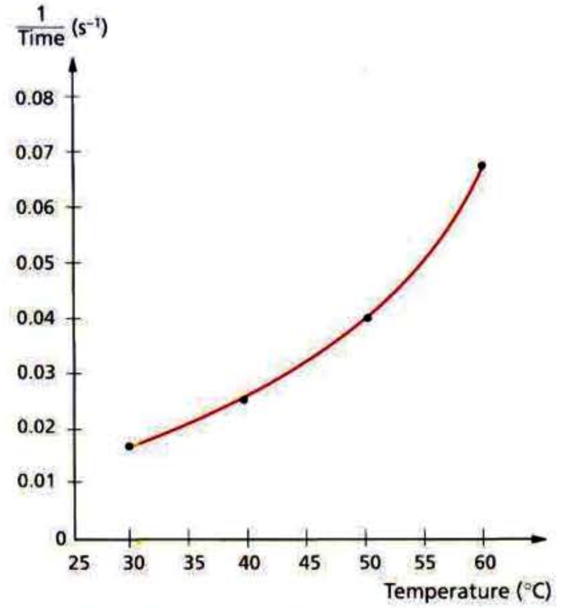


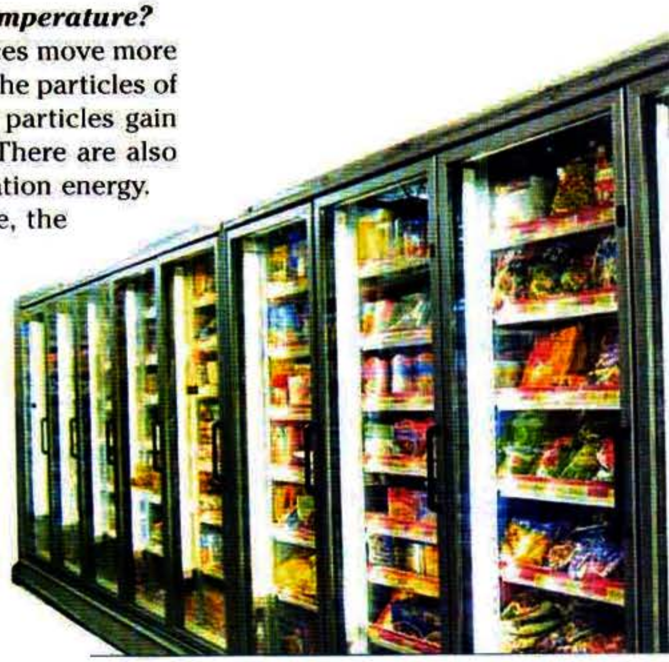
Fig. 18.20 A graph showing the speed of reaction at different temperatures

Fig. 18.20 shows the graph obtained when $\frac{1}{\text{time}}$ is plotted against temperature. The graph shows clearly *that the speed of reaction is faster at higher temperatures.*

Why does the speed of reaction increase with temperature?

At low temperatures, particles of reacting substances move more slowly because they have less energy. On heating, the particles of the reacting substances absorb energy. When the particles gain energy, they move faster and collide more often. There are also more particles with energy greater than the activation energy. Thus there are more effective collisions. Therefore, the speed of reaction increases.

This is the basis for storing fish, meat and other types of fresh foods in deep-freeze compartments where the temperature is about –20 °C. The food is kept fresh for long periods because the very low temperature slows down chemical reactions in food.



Freezing food slows down the reactions that causes food to decay.

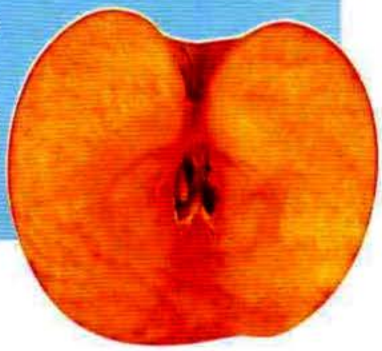
Link

Recall what you have learnt about the effect of pressure on the boiling point of a liquid in chapter 3.

This is also the reason why pressure cookers speed up cooking. The pressure cooker is a kitchen appliance that allows the pressure inside it to become greater than the atmospheric pressure. Since the boiling point of a liquid increases as the pressure above its surface increases, a pressure cooker enables water to boil at a temperature higher than 100 °C. Thus, food cooks faster in a pressure cooker.

Try it Out

1. **Apples turn brown when they are cut up and exposed to the air. In your own home, try out different methods of slowing down this browning process.**
2. **Washing powders work best at temperatures above room temperature. Think carefully about how you would measure the cleansing power of the washing powder you use at home. Then plan an experiment to determine the temperature at which your washing powder has the maximum cleansing power.**



Key Ideas

The higher the temperature, the faster the movement of the particles and the greater the number of collisions. Hence, reactions proceed faster when the temperature is increased.

Test Yourself 18.4

Questions

1. Suggest **three** ways to speed up the reaction between aluminium foil and dilute hydrochloric acid.
2. Give **two** reasons why the exhaust pipe of a car rusts more rapidly than the other metal parts of the car.

Effect of Catalysts and Enzymes on Speed of Reaction

What is a catalyst?

A catalyst is a *substance which increases the speed of a chemical reaction and remains chemically unchanged at the end of the reaction.*

What are the characteristics of a catalyst?

A catalyst has the following characteristics:

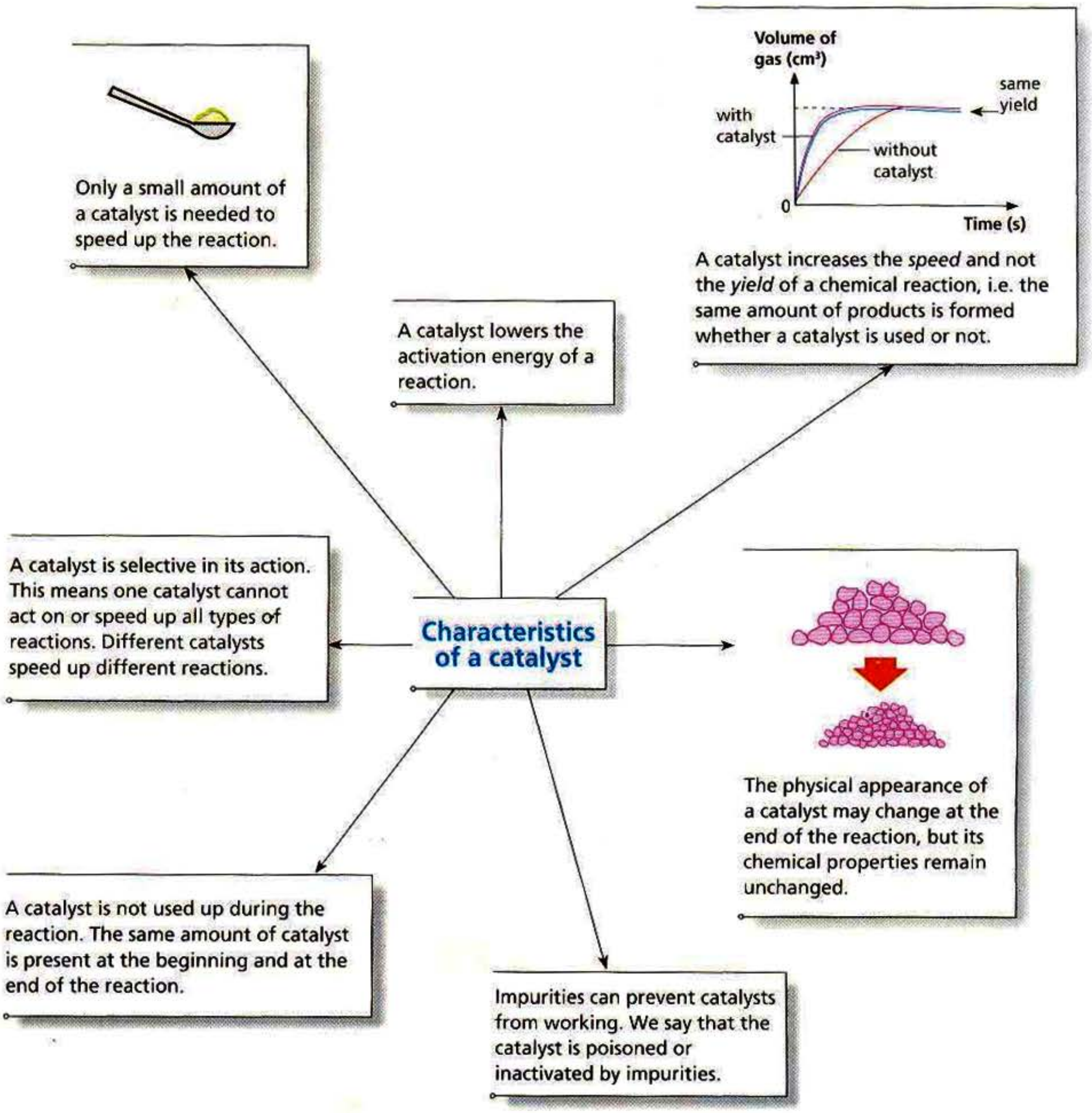


Fig. 18.21 **The characteristics of a catalyst**

How does the catalyst, manganese(IV) oxide, affect the decomposition of hydrogen peroxide?

Under normal conditions, hydrogen peroxide decomposes very slowly to give water and oxygen.

$$\text{hydrogen peroxide} \xrightarrow{\text{decomposition}} \text{water} + \text{oxygen}$$

$$2H_2O_2(aq) \xrightarrow{\text{decomposition}} 2H_2O(l) + O_2(g)$$

Let us see what happens when a very small amount of manganese(IV) oxide is added to hydrogen peroxide.

Experiment 6

To study the effect of manganese(IV) oxide on the speed of decomposition of hydrogen peroxide solution

Procedure

1. Two test tubes containing 10 cm^3 of hydrogen peroxide solution are set up.
2. 1.0 g of black manganese(IV) oxide is added to the solution in one test tube.
3. A glowing splint is placed at the mouth of each test tube and the observations are recorded.

In the test tube where manganese(IV) oxide is added to hydrogen peroxide, bubbles of oxygen gas are quickly produced. The glowing splint is rekindled and burns brightly.

This experiment shows that manganese(IV) oxide speeds up the decomposition of hydrogen peroxide. It is called a *catalyst* for the reaction.

However, manganese(IV) oxide is not the only catalyst that can be used to speed up the decomposition of hydrogen peroxide. Other suitable catalysts are copper(II) oxide and iron(III) hydroxide.

How would you compare the effectiveness of different catalysts?

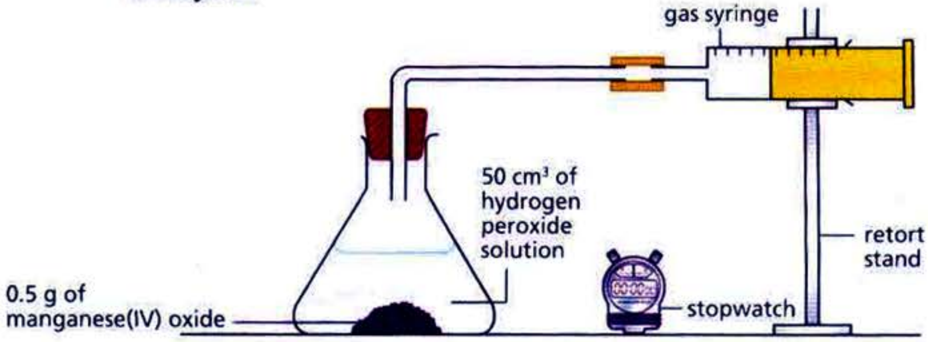


Fig. 18.22 An experiment to study the effectiveness of different catalysts

We can use the apparatus shown in Fig. 18.22 to compare the effectiveness of copper(II) oxide and manganese(IV) oxide as catalysts for the decomposition of hydrogen peroxide.

- In experiment I, 0.5 g of powdered manganese(IV) oxide is used to catalyse the decomposition of 50 cm^3 of hydrogen peroxide.
- In experiment II, 0.5 g of powdered copper(II) oxide is used to catalyse the decomposition of the same amount of hydrogen peroxide.

Powdered manganese(IV) oxide is used to catalyse the decomposition of hydrogen peroxide.

The results of the experiments are shown in Fig. 18.23.

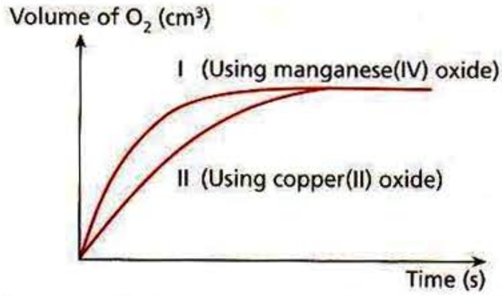


Fig. 18.23 Comparing the effectiveness of manganese(IV) oxide and copper(II) oxide in the decomposition of hydrogen peroxide

Graph I is steeper than graph II. This shows that the reaction is faster using manganese(IV) oxide as the catalyst. Thus, manganese(IV) oxide is a better catalyst than copper(II) oxide.

What are the common catalysts for some industrial processes?
In many industrial processes, catalysts are used to speed up the various reactions. Below are some examples:

- Iron in the Haber process for manufacturing ammonia (chapter 19)
- Vanadium(V) oxide in the manufacture of sulphuric acid
- Platinum or rhodium in catalytic converters (chapter 20)
- Aluminium oxide or silicon(IV) oxide in the cracking process (chapter 23) for producing hydrogen

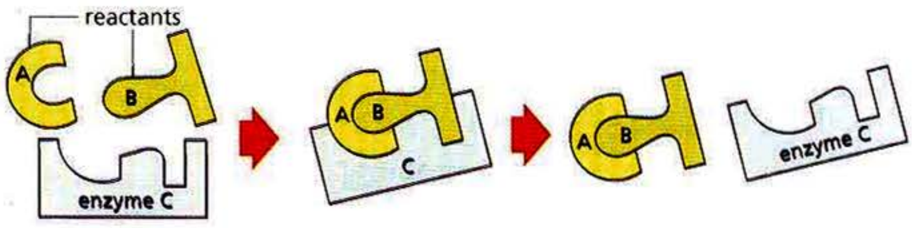
Chem-Aid

1. Catalysts are often used in the form of pellets or gauzes. This makes the catalyst more effective because of the larger surface area.
2. Some substances slow down the speed of chemical reactions. They are called **inhibitors.**

Enzymes — biological catalysts
Enzymes are *substances that catalyse the chemical reactions in plants and animals.* They are often called **biological catalysts.** Enzymes have the following properties:

1. *They are proteins.*
2. *They are very specific in their actions.* The human body contains many different kinds of enzymes and each one does a single specific job. Amylase, for example, is the enzyme present in saliva. It catalyses the reaction which changes starch into sugar.

3. *They can be made inactive by heating.* Many enzymes operate most effectively at body temperature, i.e. between 35 °C and 40 °C. Above or below this range, the speed of reaction decreases. Too low a temperature will render enzymes inactive while too high a temperature will cause enzymes to become *denatured* (i.e. destroyed).
4. *They are sensitive to pH changes.* Most enzymes have a certain range of pH at which they work best.



reactant A + reactant B $\xrightarrow{\text{enzyme C}}$ product AB

Fig. 18.24 **Enzymes contain reactive sites. Each reactive site only fits a particular molecule.**

What are the uses of enzymes in industry?

Enzymes are important in industry. In the manufacture of wine and other alcoholic drinks, enzymes produced by yeast are used to catalyse the conversion of sugar or starch to ethanol.

sugar or starch $\xrightarrow{\text{yeast}}$ ethanol + carbon dioxide

Modern detergents (biological washing powders) contain enzymes which are very effective in removing stains caused by protein-based substances such as food and blood.

Catalysts and activation energies

Catalysts work by providing an alternative pathway for the reaction to proceed (reaction pathway). This reaction pathway has a lower activation energy than the reaction pathway for the original reaction. The effect of catalysts on activation energies can be shown by energy profile diagrams such as those in Fig. 18.25.

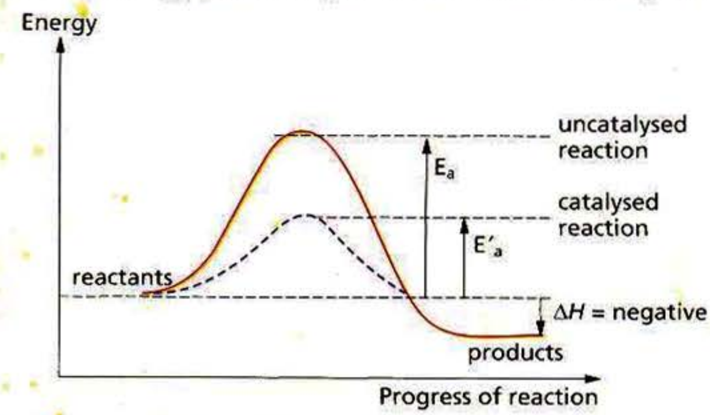


Fig. 18.25 **Energy profile diagrams showing the effect of a catalyst on activation energy**

Study Fig. 18.25. You can see that the activation energy of the uncatalysed reaction is E_a and that of the catalysed reaction is E'_a. E'_a is less than E_a. Since the activation energy is lower for the catalysed reaction, a greater proportion of collisions between reacting particles will result in product particles being formed much faster. Hence, the speed of reaction will be greater for the catalysed reaction.

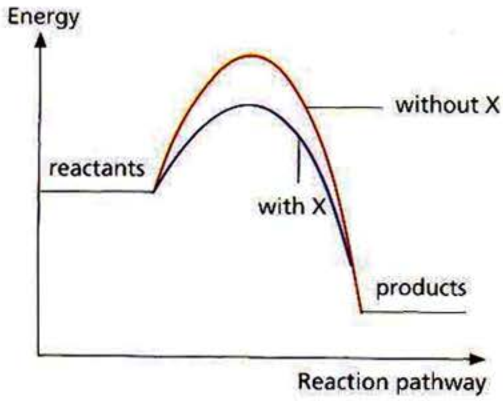
Key ideas

1. A catalyst is a substance which increases the speed of a chemical reaction.
2. A catalyst increases the speed but not the yield of a chemical reaction.
3. The mass and chemical properties of a catalyst remain unchanged at the end of the reaction.
4. Enzymes are found in living things. They are biological catalysts. Many reactions which occur in living cells are brought about by enzymes.
5. A catalyst increases the speed of a reaction by providing an alternative reaction pathway with a lower activation energy.

Test Yourself 18.5

Worked Example

The energy profile diagrams for a reaction without X and in the presence of X are shown below.



a) What is the effect of X on
   i) the enthalpy of the reaction?
   ii) the rate of reaction?
b) What is the role of X in this reaction?

Thought Process

The enthalpy of the reaction is the difference between the energy content of the reactants and the products.

If the activation energy is lowered, the reaction rate increases. In the presence of X, the activation energy is lowered. Hence, the rate of reaction is higher.

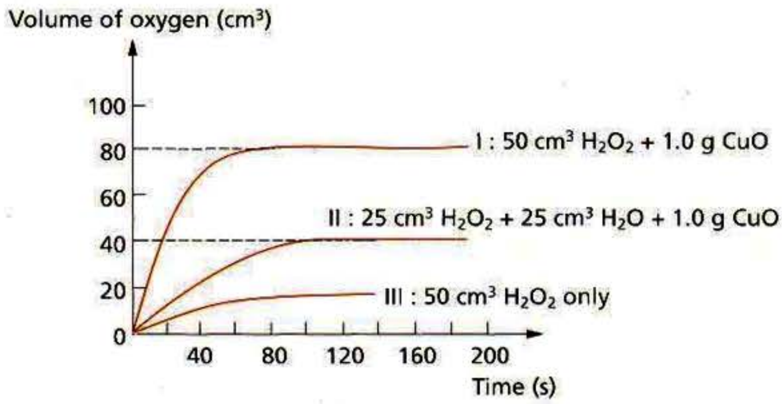
Answer

a) i) The enthalpy of the reaction remains unchanged.
   ii) The rate of reaction increases as the activation energy is lowered.
b) X acts as a catalyst in this reaction.

Questions

1. The decomposition of hydrogen peroxide is catalysed by manganese(IV) oxide. How would you show that the catalyst is not used up at the end of the reaction?

2. Graphs I, II and III below show the results obtained when hydrogen peroxide is decomposed under different conditions.



   a) Which graph represents
      i) the slowest reaction?
      ii) or the fastest reaction?
   b) Explain why
      i) the reaction rate under condition I differs from that under condition II, and
      ii) the reaction rate under condition II differs from that under condition III.

Concept Map

Chemical Reaction

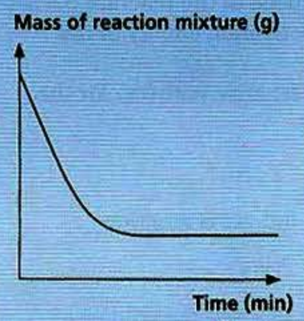
occurs only if

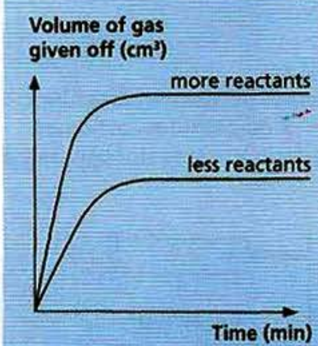
- Reacting particles collide with one another
- Colliding particles have activation energy

the rate at which it proceeds is called

Speed of reaction

shown by these typical graphs





can be measured from

Change in mass of reactant/product / Time taken — represented by gradient of graph

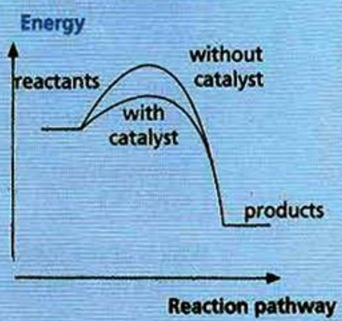
or

Change in volume of gas given off / Time taken — represented by gradient of graph

can be increased by

using **Catalysts** which

Lower activation energy → more particles have sufficient energy to collide to form products



increasing

Concentration of reactants since More reacting particles occupying the same volume → more collisions

Temperature of reactants since Colliding particles have more energy → more collisions

Pressure (for gases) since Reacting particles are closer → more collisions

decreasing

Particle size of solid reactants since Greater surface area for reacting particles to collide → more reactions

Exercise 18

Foundation

1. Calcium carbonate reacts with dilute hydrochloric acid according to the equation:

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

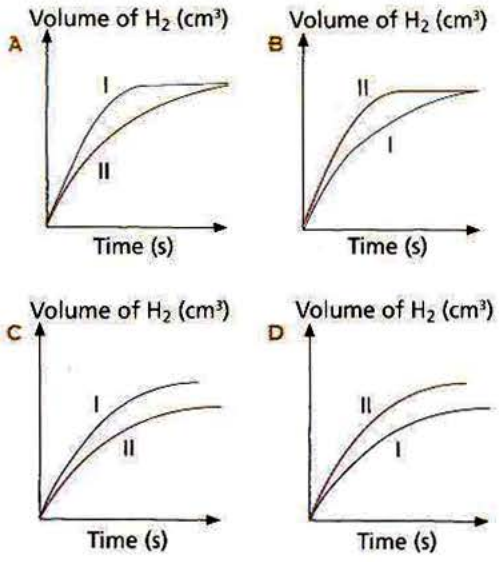
The rate of reaction decreases with time because the ______________ .

A concentration of the product ($CaCl_2$) decreases
B concentration of hydrochloric acid decreases
C temperature of hydrochloric acid decreases
D volume of hydrochloric acid decreases

2. Two experiments were carried out to study the speed of reaction between zinc and dilute sulphuric acid under the following conditions.

Experiment	I	II
Mass of zinc powder (g)	1.0	1.0
Volume of sulphuric acid (cm^3)	20	20
Concentration of sulphuric acid (mol/dm^3)	0.1	0.1
Temperature of sulphuric acid (°C)	28	35

Which of the following graphs represent the results for experiments I and II?



3. Under which of the following conditions will marble (calcium carbonate) react most rapidly with hydrochloric acid?

A Marble chips and dilute acid at 40 °C.
B Marble chips and concentrated acid at 20 °C.
C Marble powder and dilute acid at 20 °C.
D Marble powder and concentrated acid at 40 °C.

4. Which of the following statements about the effect of a catalyst is correct?

A It increases the yield of the product.
B It increases the speed of the reactant particles.
C It increases the activation energy of the reaction.
D It provides an alternative pathway for the reaction to occur.

5. Explain the following.

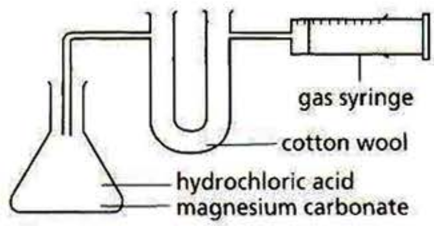
a) The speed of the following reaction

$$H_2(g) + I_2(g) \longrightarrow 2HI(g)$$

is found to increase when the volume of the container is decreased.

b) Exposed steel rusts extremely slowly in the Antarctic.

6. The speed of the reaction between hydrochloric acid and large pieces of magnesium carbonate (in excess) was investigated using the apparatus below.



50 cm^3 of hydrochloric acid was added to 20 g of magnesium carbonate and the total volume of gas given off was recorded every ten seconds as shown in the table below.

Time (s)	0	10	20	30	40	50	60	70
Total volume of gas given off (cm^3)	0	3	12	27	49	68	86	99

a) Plot the results of the experiment on graph paper.

b) Sketch, on the same axes, the curves you would expect to obtain if the experiment was repeated using:
   i) 25 cm^3 of hydrochloric acid of the same concentration.
   ii) 20 g of smaller pieces of magnesium carbonate.

c) Explain your answer in (b)(ii) based on the collision of particles.

d) Why was the reaction slow at the beginning?

e) Explain the purpose of using the cotton wool in the apparatus set-up.

7. 3.0 g of marble chips are added to an excess of 2.0 mol/dm^3 dilute hydrochloric acid for the following reaction to occur.

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

You are asked to follow the speed of this reaction by measuring the mass lost from the reaction system.

a) Draw a labelled diagram to show how the experiment can be carried out.

b) Sketch a graph of total mass loss against time to show the results you would expect to obtain in this experiment.

c) Based on the gradient of the graph you have drawn, what can you infer about the speed of the reaction as the reaction proceeds?

d) Explain your answer in (c) in terms of collisions between the reacting particles.

e) The experiment was repeated using 6.0 g of marble chips and an excess of 2.0 mol/dm^3 hydrochloric acid. Draw the curve that you would expect to get for this experiment on the axes in (b).

8. Three experiments were carried out to study the speeds of reaction between dilute acids and granulated zinc. The amount of chemicals used and the time taken to collect 30 cm^3 of hydrogen gas were recorded for each experiment as shown in the table on the right.

Experiment	Amount of dilute acid used	Amount of granulated zinc used (g)	Amount of copper(II) sulphate	Time taken (s)
I	25.0 cm^3 of 1.0 mol/dm^3 H_2SO_4	2.5	0	28.0
II	25.0 cm^3 of 1.0 mol/dm^3 HCl	2.5	0	56.0
III	25.0 cm^3 of 1.0 mol/dm^3 HCl	2.5	5 drops	14.0

a) Write the ionic equation for the reaction between zinc and the dilute acids.

b) What is the average speed of reaction for the first 28 s in experiment I?

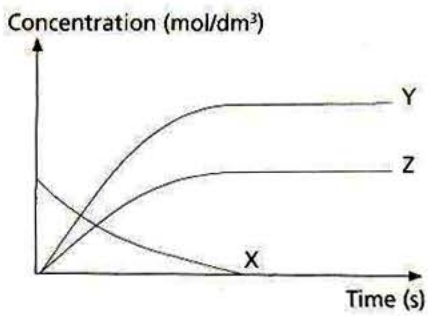
c) Explain why the time taken to collect 30 cm^3 of hydrogen gas in experiment I is shorter than in experiment II.

d) Explain why the time taken to collect the hydrogen gas in experiment III is shorter than in experiment II.

e) Using the same axes, draw the energy profile diagram for experiments II and III.

Challenge

1. The following diagram shows the change in reactant and product concentrations with time during a chemical reaction.



The chemical reaction represented is ______.

A X $\longrightarrow$ Y + Z **B** X $\longrightarrow$ 2Y + Z
C Z $\longrightarrow$ 2X + Y **D** Z $\longrightarrow$ 2Y + X

2. If a piece of zinc is added to excess sulphuric acid, an exothermic reaction occurs. The rate of reaction increases during the first few seconds because ________.

A the reaction mixture is getting hotter
B the amount of zinc is decreasing
C the Zn^{2+} ions produced act as a catalyst
D the concentration of sulphuric acid is increasing

3. Sodium thiosulphate reacts with hydrochloric acid according to the equation

$$Na_2S_2O_3(aq) + 2HCl(aq) \longrightarrow S(s) + 2NaCl(aq) + SO_2(g) + H_2O(l)$$

This reaction can be used to investigate the effect of concentration on the speed of reaction between sodium thiosulphate and hydrochloric acid. The table below shows the results of five experiments. In each experiment, different volumes of sodium thiosulphate were added to 10 cm^3 of hydrochloric acid. The time taken for the precipitate to form was measured. All the experiments were carried out at 25 °C.

Experiment	Volume of sodium thiosulphate (cm^3)	Volume of water (cm^3)	Time taken for the precipitate to form (s)
1	50	0	43
2	40	10	55
3	30	20	75
4	20	30	124
5	10	40	255

a) What conclusion can you make from these results?

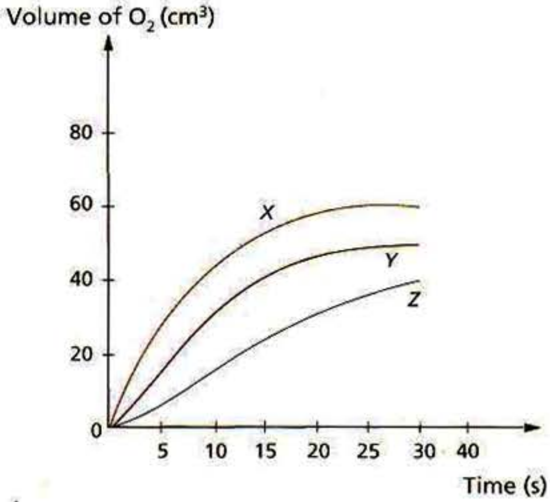
b) i) Plot the results on graph paper to show how the volume of sodium thiosulphate solution affects the reaction time.
ii) Sketch, on the same graph paper, the graph you would expect to obtain if the experiments were repeated at 40 °C.

c) Explain your answers in (a) and (b)(ii) separately, based on the collision of particles.

4. Three experiments were carried out to compare the effects of three different catalysts, *X*, *Y* and *Z*, on the decomposition of hydrogen peroxide.

$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$

In each experiment, the same amount of catalyst was added to 20 cm^3 of 0.25 mol/dm^3 hydrogen peroxide. The results of the experiments are shown below.



a) Name **three** factors that must be kept constant when carrying out the experiments.

b) Which is the best catalyst for this experiment?

c) i) Predict the final volume of oxygen produced using *Z* as the catalyst.
ii) Explain your answer.

d) One of the catalysts used was copper(II) oxide. Describe what was left at the end of the reaction.

e) Another catalyst used was manganese(IV) oxide. Outline the steps that can be included in the experiment to show that all the oxygen produced is from hydrogen peroxide alone, not from manganese(IV) oxide.

Chemistry Today

Chemical reactions are taking place all the time inside living organisms. These reactions keep the organisms alive and active. The rate at which the reactions happen is called the metabolic rate of an organism.

Alligators are cold-blooded animals. Did you know that their metabolic rates depend on the temperature of their surroundings? In the evening, alligators usually become sleepy and inactive because a drop in temperature slows down their metabolic rates. In the morning, alligators bask in the sun to increase their body temperature. This increases their metabolic rates. Without warming up this way, the alligators would stay sleepy and inactive all day!

CRITICAL THINKING

Suggest a reason why humans can maintain a rather constant metabolic rate that is not dependent on the surrounding temperature.

Chapter 19

Ammonia

Chapter Outline

Ammonia (NH_3) is an extremely useful chemical. It is used to make explosives. It is also used to make fertilisers, which are extremely important in agriculture. It can even be used to make cleansers. However, for many years, there was no efficient way to produce ammonia in large quantities until the Haber process was invented by Fritz Haber. In this chapter, you will learn about ammonia and the Haber process.

19.1 | Reversible Reactions

Many chemical reactions can proceed in one direction only, i.e. they cannot be reversed. For example, potassium hydroxide reacts with dilute sulphuric acid to form potassium sulphate and water:

$$2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$$

However, there are some chemical reactions that can be reversed. For example, when vapour of concentrated ammonia solution comes into contact with vapour of concentrated hydrochloric acid, white fumes of ammonium chloride are formed. The equation for this reaction is

$$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$$

Solid ammonium chloride, in turn, can be decomposed on heating to form ammonia and hydrogen chloride gases. The equation for this reaction is

$$NH_4Cl(s) \longrightarrow NH_3(g) + HCl(g)$$

Since *the reaction can go in either direction*, we say that it is a **reversible reaction**. A double arrow sign, $\rightleftharpoons$, is used to indicate a reversible reaction. Therefore, the equation for the reversible reaction of ammonia and hydrogen chloride should be written as

$$NH_3(g) + HCl(g) \rightleftharpoons NH_4Cl(s)$$

There are many other examples of reversible reactions. One such example is the reaction between nitrogen and hydrogen to form ammonia. It is written as

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

To avoid confusion, chemists always call
- the reaction from left to right, the **forward reaction**,
- the reaction from right to left, the **reverse reaction**.

Reversible reaction

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Forward reaction	Reverse reaction
$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$	$2NH_3(g) \longrightarrow N_2(g) + 3H_2(g)$
reactants — product	reactant — products

Experiment 1

To study the reversible reaction of potassium chromate(VI)

Procedure

1. Place 2 cm^3 of potassium chromate(VI) solution in a boiling tube. Note the colour of the solution.
2. Add dilute sulphuric acid slowly. Stop adding acid when the solution changes colour. Note the colour of the solution.
3. Add aqueous sodium hydroxide slowly to the resulting solution obtained in step 2. Stop adding sodium hydroxide when the solution changes colour. Note the colour of the solution.

Questions

1. What is the colour of potassium chromate(VI) in the presence of (a) an acid and (b) an alkali?
2. Write an ionic equation for the reversible reaction that occurred in this experiment.

19.2 Manufacturing Ammonia by the Haber Process

The raw materials for the manufacture of ammonia in the Haber process are nitrogen and hydrogen. Nitrogen is extracted from air (see chapter 20) and hydrogen is produced from the cracking of petroleum (see chapter 23).

Try it Out

Surf the Internet and read up more about Fritz Haber, who invented the Haber process. Find out why he got involved in the manufacture of ammonia gas

Since the nitrogen molecule is unreactive, there is no reaction at all between nitrogen and hydrogen at room temperature and pressure. To make nitrogen react with hydrogen to form ammonia, a high pressure and a relatively high temperature are needed. Iron is used to speed up the reaction, i.e. as a catalyst. The equation for the Haber process is

$$\text{nitrogen} + \text{hydrogen} \xrightleftharpoons{\text{finely divided iron}} \text{ammonia}$$

Conditions Required for Manufacturing Ammonia

The reaction between hydrogen and nitrogen to form ammonia is a reversible process. This means that some of the ammonia formed may revert back to nitrogen and hydrogen. So to achieve the maximum yield of ammonia at the minimum cost, the reaction conditions are very carefully controlled.

Link

The speed of any reaction can be controlled by controlling conditions like pressure and temperature. Recall what you have learnt in chapter 18.

1 What is the effect of temperature on the yield of ammonia?
Fig. 19.1 shows the yield of ammonia at different temperatures and pressures. We can see that a lower temperature increases the yield of ammonia and reduces the decomposition of ammonia to hydrogen and nitrogen. However, a lower temperature also results in a slow reaction.

2 What is the effect of pressure on the yield of ammonia?
The yield of ammonia is increased under higher pressures. High pressure also increases the rate of reaction. However, maintaining high pressure is costly because expensive equipment (for example, special pumps and stronger pipes) are required.

3 What is the effect of a catalyst on the yield of ammonia?
A high pressure and a relatively high temperature are needed to make nitrogen react with hydrogen to form ammonia. Even then the reaction is slow, so a catalyst is used to speed up the reaction.

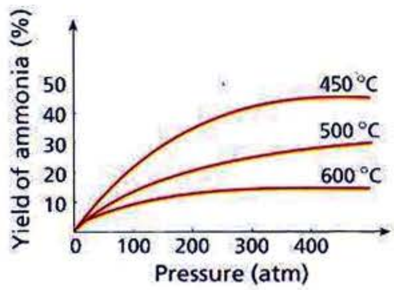


Fig. 19.1 Graph of yield of ammonia at different temperatures and pressures

Quick Check

In the reaction between nitrogen and hydrogen, more ammonia can be produced if the pressure on the mixture of gases is increased. Give two reasons why most factories do not use pressures higher than 300 atm in the manufacture of ammonia.

What are the operating conditions in the Haber process?

It was found that the best conditions for producing ammonia in the Haber process are

- a temperature of 450 °C,
- a pressure of 250 atm, and
- the presence of iron catalyst.

Fig. 19.2 shows the flow diagram for the manufacture of ammonia by the Haber process.

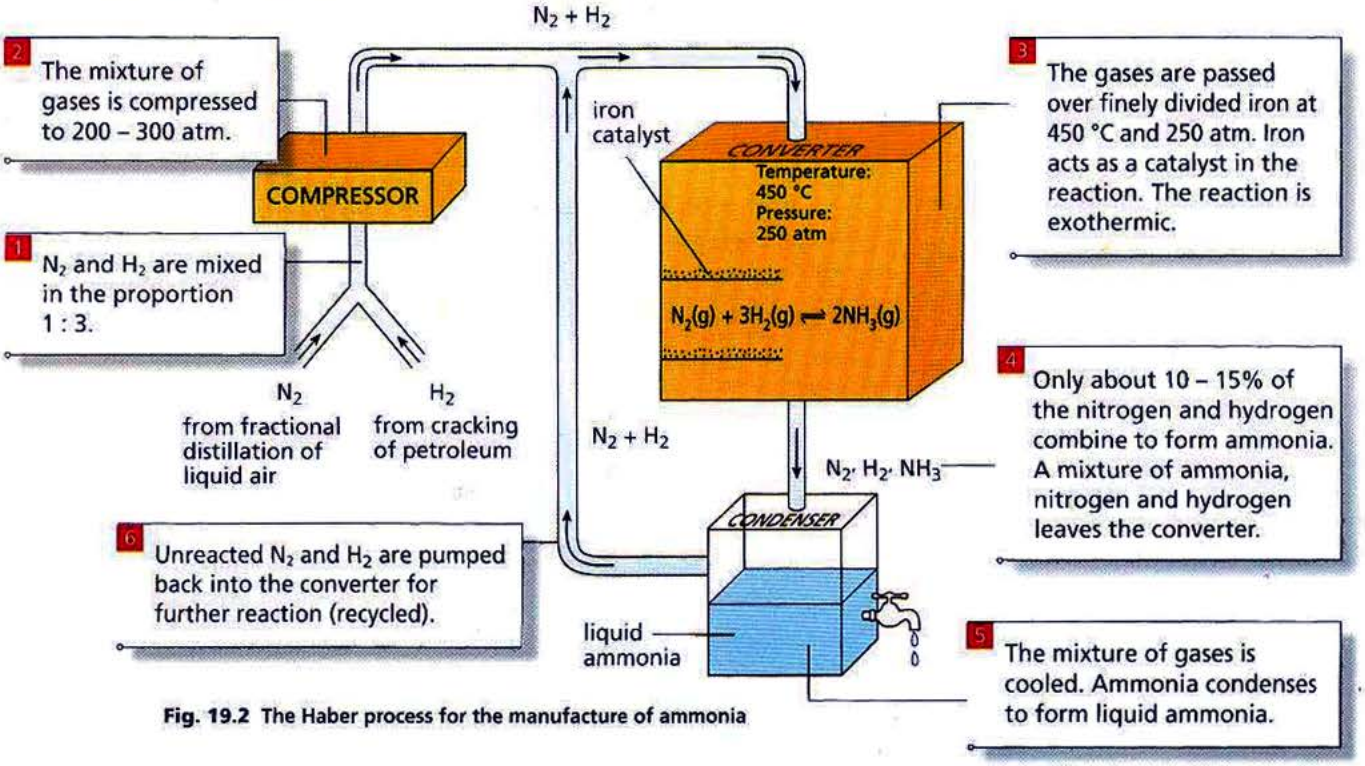


Fig. 19.2 The Haber process for the manufacture of ammonia

19.3 | Displacement of Ammonia from its Salts

Look at the two reactions below:

ammonium chloride + sodium hydroxide $\xrightarrow{\text{heat}}$ ammonia gas + water + sodium chloride

$$NH_4Cl(s) + NaOH(aq) \xrightarrow{\text{heat}} NH_3(g) + H_2O(l) + NaCl(aq)$$

ammonium chloride + calcium hydroxide $\xrightarrow{\text{heat}}$ ammonia gas + water + calcium chloride

$$2NH_4Cl(s) + Ca(OH)_2(aq) \xrightarrow{\text{heat}} 2NH_3(g) + 2H_2O(l) + CaCl_2(aq)$$

In both cases, you can see that ammonia gas is produced on heating. Whenever an ammonium salt is heated with an alkali (for example, sodium hydroxide or calcium hydroxide), ammonia is displaced from the salt.

TidBit

Calcium hydroxide and calcium oxide (quicklime) are used widely in agriculture to neutralise excess acidity in soil. However, calcium hydroxide has one undesirable effect when used in soil treatment. What do you think is the undesirable effect? Both calcium hydroxide and calcium oxide react with nitrogenous fertilisers to form ammonia gas, which then escapes into the atmosphere. This causes the loss of nitrogen from fertilisers already added to the soil by farmers. Find out why nitrogen is important to plants on page 358.

Key ideas

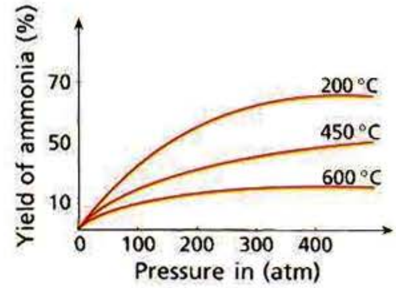
1. A reversible reaction is a reaction that can take place in both directions, i.e. from left to right as well as from right to left of the equation.
2. The reaction between nitrogen and hydrogen to form ammonia is a reversible reaction.
3. The Haber process is used to manufacture ammonia.
4. In the Haber process, nitrogen and hydrogen react to produce the highest yield of ammonia under the following conditions:
   - Temperature: 450 °C
   - Pressure: 250 atm
   - Catalyst: finely divided iron
5. Ammonia is displaced from its salt when it is heated with an alkali.

Farmers add nitrogenous fertilisers to soil to make crops grow better.

Test Yourself 19.1

Worked Example 1

The graph shows the yield of ammonia at different conditions. Explain why a temperature of 450 °C is used in the Haber process instead of 200 °C and 600 °C.



At 200 °C, the rate of production of ammonia is too low. At 600 °C, the yield of ammonia is too low as ammonia is decomposed into its reactants. Hence, 450 °C is the best temperature for the process to operate at.

Worked Example 2

Calculate the nitrogen content of ammonium nitrate.

Answer

Relative molecular mass of ammonium nitrate, NH_4NO_3,
$= 14 + (4 \times 1) + 14 + (3 \times 16) = 80$

Mass of nitrogen in one mole of the salt
$= 2 \times 14 = 28$ g

Percentage of nitrogen in ammonium nitrate
= nitrogen content (%)

$$= \frac{\text{mass of nitrogen in the salt}}{\text{relative formula mass of the salt}} \times 100\%$$

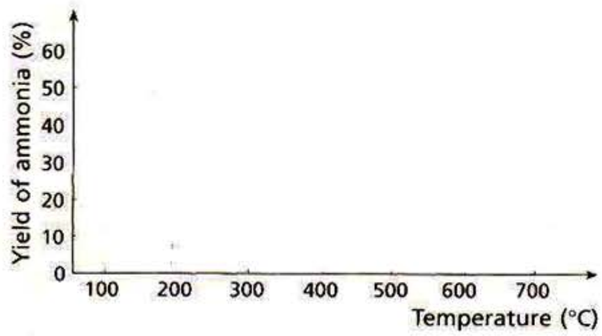
$$= \frac{28}{80} \times 100\%$$

$= 35\%$

Questions

1. Write a chemical equation for each of the following reversible reactions.
   a) Sulphur dioxide reacts with oxygen to form sulphur trioxide.
   b) Hydrogen gas and iodine gas react to form gaseous hydrogen iodide.

2. a) Copy the axes on the right and sketch the graphs that you would expect to get if the reactions between nitrogen and hydrogen were carried out at a pressure of (i) 500 atm and (ii) 300 atm. Use the same axes for both graphs.



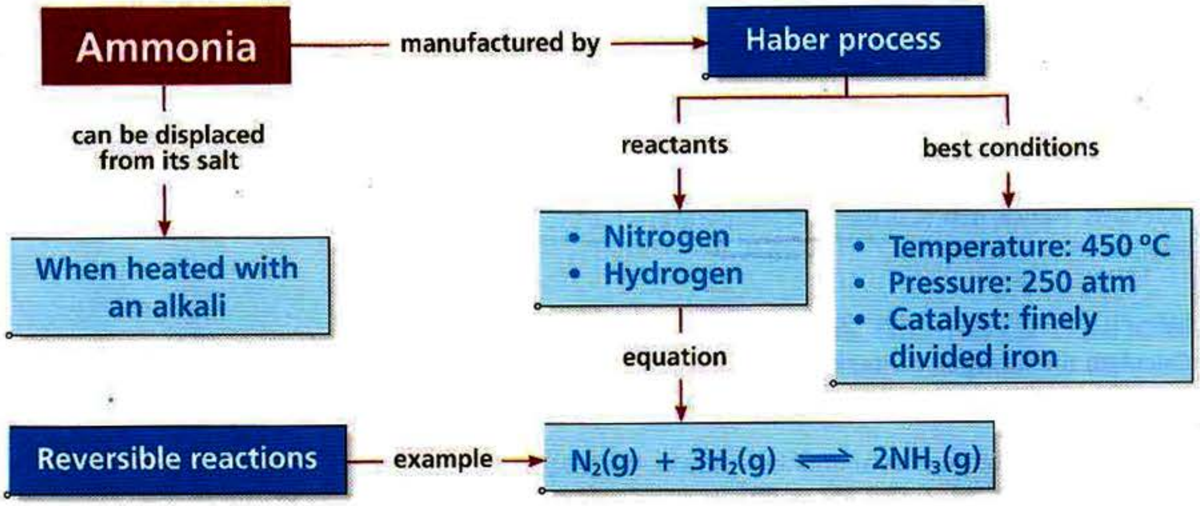
   b) Comment on the graphs that you have drawn.

3. What is the nitrogen content of urea, $CO(NH_2)_2$?

Try it Out

Do you know what 'Blue Baby Syndrome' is? Do a research project on this illness. Visit the World Health Organization's (WHO) website at http://www.who.int/water_sanitation_health/diseases/methaemoglob/en/ to learn more about what causes this illness and how to prevent it. Your write-up should include the symptoms of the illness, its treatment and the laws laid down by agricultural countries to reduce the incidence of the illness.

Concept Map



Exercise 19

Foundation

1. Which conditions are ideal for the manufacture of ammonia in the Haber process?

	Pressure	Temperature	Ratio of $H_2 : N_2$
A	250 atm	450 °C	3 : 1
B	250 atm	450 °C	1 : 3
C	450 atm	250 °C	3 : 1
D	450 atm	250 °C	1 : 3

2. Ammonia is made by the reaction between nitrogen and hydrogen:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

This reaction is slow at r.t.p. An increase in the yield of ammonia and the rate of reaction can be achieved by ______________.

A adding a suitable catalyst
B cooling the reactants
C increasing the pressure
D increasing the temperature

3. To prepare ammonia, ammonium chloride should be heated ______________.

A on its own
B with water
C with nitric acid
D with sodium hydroxide

4. Which of the following reactants will not produce ammonia on heating?

A Ammonium chloride and calcium hydroxide.
B Ammonium sulphate and calcium oxide.
C Potassium hydroxide and sodium nitrate.
D Sodium hydroxide and ammonium nitrate.

5. Farmers usually treat soil with ammonium sulphate because it ______________.

A increases the nitrogen content of the soil
B increases the sulphate content of the soil
C reduces the acidity of the soil
D reduces the alkalinity of the soil

6. Some fertilisers have the label NPK on them to indicate that they contain the elements nitrogen, phosphorus and potassium. Which of the following mixtures of compounds cannot be sold as an NPK fertiliser?

A Ammonium chloride and potassium hydrogen phosphate.
B Ammonium sulphate and potassium chloride.
C Potassium hydrogen sulphate and ammonium hydrogen phosphate.
D Sodium phosphate and potassium nitrate.

Challenge

1. a) What is meant by 'a reversible reaction'?

   b) When hydrated copper(II) sulphate is heated, anhydrous copper(II) sulphate is formed.

      i) Describe what is observed during the heating process.

      ii) What does the following equation show?

      $$CuSO_4.5H_2O \rightleftharpoons CuSO_4 + 5H_2O$$

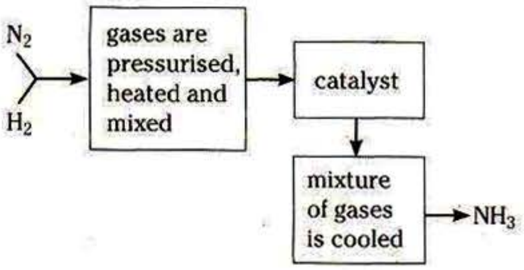
      iii) Devise an experiment to prove that hydrated copper(II) sulphate can be converted to anhydrous copper(II) sulphate and vice versa.

   c) The reaction between nitrogen and hydrogen is often carried out in the presence of finely divided iron.

      i) Is this a redox reaction? Explain your answer.

      ii) What effect does iron have on this reaction?

2. The flow chart for making ammonia is shown below.



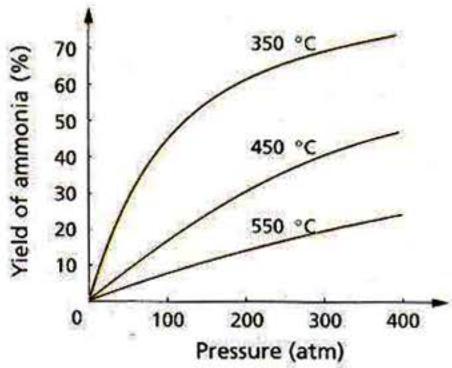
   a) Write the balanced equation for the reaction between nitrogen and hydrogen to form ammonia.

   b) i) After passing over the catalyst, the mixture of gases contains nitrogen, hydrogen and ammonia. Explain why hydrogen and nitrogen are present in the mixture.

      ii) When this mixture of gases is cooled, only ammonia turns into a liquid. Explain this observation.

      iii) Explain what should be done to the hydrogen and nitrogen in the mixture of gases.

3. Large quantities of ammonia are manufactured by the Haber process. The graph below shows the yield of ammonia at different temperatures and pressures.



   a) Deduce from the graph the best conditions for obtaining a good yield of ammonia.

   b) Suggest why these conditions are not necessarily used by industries to manufacture ammonia.

   c) If 24 dm^3 of nitrogen reacted completely with 72 dm^3 of hydrogen, what volume of ammonia would be formed?
   (All volumes were measured under the same conditions.)

   d) In one industrial process, only 14.4 dm^3 of ammonia was obtained by using the volumes described in (c). Suggest one method of increasing the yield of ammonia without changing the pressure and temperature of the reaction.

4. a) Write the balanced equation for the reaction between ammonium chloride and calcium hydroxide.

   b) Describe a test for nitrate ions in aqueous solution.

   c) Describe a test for nitrate ions in an aqueous solution that also contains ammonium ions.

   d) A farmer wants to neutralise as well as fertilise his farmland. He decided to add both calcium hydroxide and ammonium nitrate to the soil at the same time. Is this a good decision? Why?

Chemistry Today

Nitrogen is a vital source of food for plants. It is required in large amounts by plants to make plant proteins. These proteins are necessary for the growth and repair of plant cells. Thus, nitrogen is used widely in the manufacture of fertilisers to promote plant growth and increase crop yields.

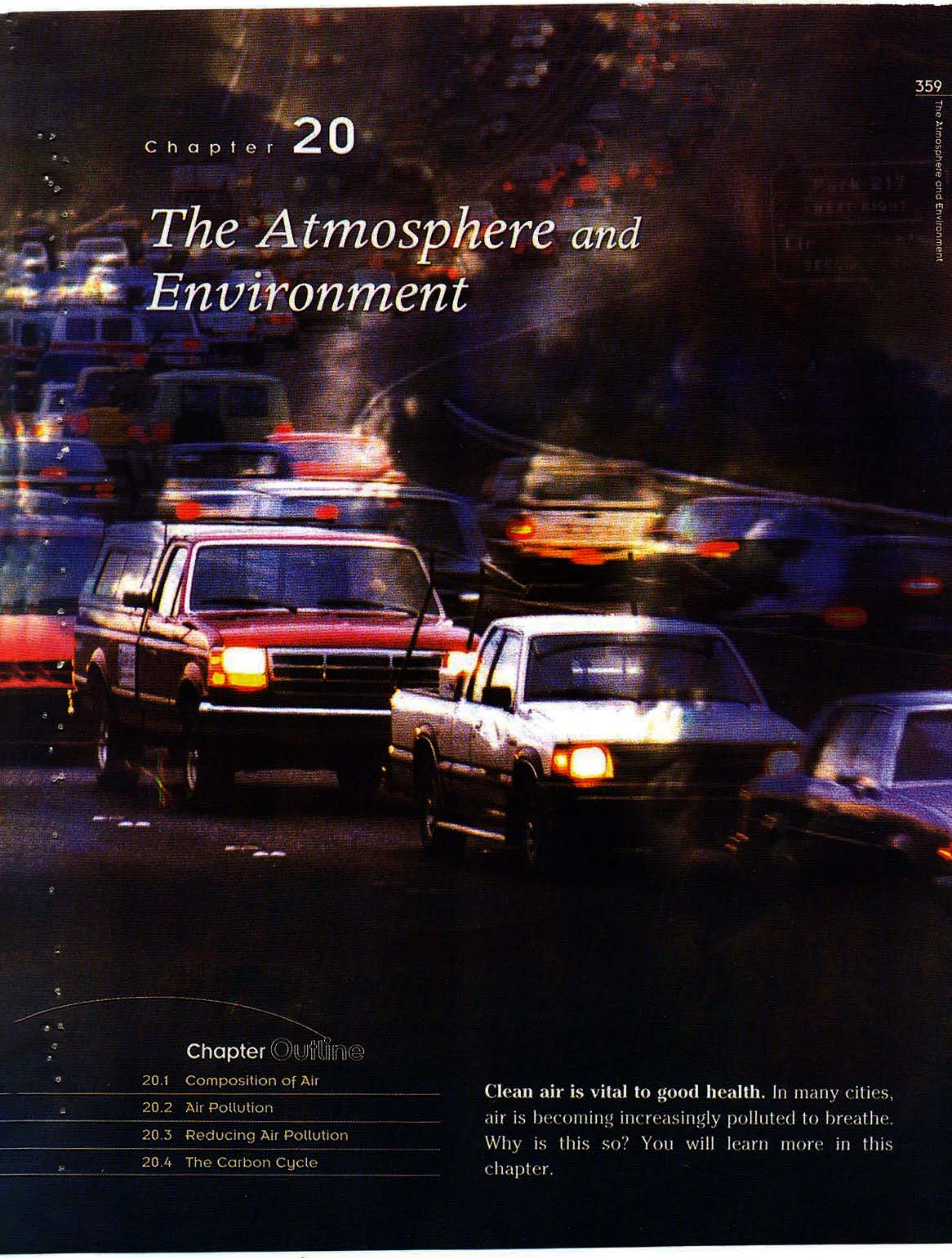
Although there is plenty of nitrogen in the air (78 – 79%), many plants cannot make use of atmospheric nitrogen directly because they cannot absorb the gas from the air. Only plants with special nodules on their roots can do this.

Most plants get their nitrogen supply by absorbing soluble nitrogen compounds from the soil. Hence, nitrates, ammonia and ammonium salts are used as fertilisers. Fertilisers that contain nitrogen are called nitrogenous fertilisers. Some nitrogenous fertilisers are shown below.

Nitrogenous fertiliser	Formula	Comment
Liquid ammonia	NH_3	• injected into the ground using special equipment
Ammonium nitrate	NH_4NO_3	• made from ammonia and nitric acid
Ammonium sulphate	$(NH_4)_2SO_4$	• made from ammonia and sulphuric acid
Urea	$CO(NH_2)_2$	• made from ammonia and carbon dioxide • less soluble in water than other ammonium fertilisers • reacts with water in the soil to form ammonia

CRITICAL THINKING

Try calculating the percentage of nitrogen by mass in each fertiliser. By comparing the values, decide which is the best fertiliser.



Chapter 20

The Atmosphere and Environment

Chapter Outline

Clean air is vital to good health. In many cities, air is becoming increasingly polluted to breathe. Why is this so? You will learn more in this chapter.

Joseph Priestley (1733 – 1804)

The British scientist, Joseph Priestley, was the first to isolate oxygen. He also discovered that the gas released by fermenting grains gave a pleasant sweet taste when dissolved in water. This gas is actually carbon dioxide. Priestley made the first carbonated soft drink! Priestley never attended a university. However, his passion for science led him to carry out many experiments that helped lay the foundation of chemistry.

20.1 | Composition of Air

The Earth is surrounded by a blanket of air which we call the **atmosphere**. Air is an important natural resource. Not only do we use oxygen in the air to breathe, we also extract many gases from the air to use as raw materials in industry.

What does air consist of?

Air is a mixture of several gases. It contains elements and compounds that are needed by all living things. As air is a mixture, its composition varies from time to time and from place to place. Table 20.1 shows the composition by volume of a typical sample of clean air.

Gas	Composition by volume
nitrogen	78 – 79%
oxygen	20%
carbon dioxide	0.03%
water vapour	0 – 5%
noble gases:	
• argon	0.9%
• neon and helium	0.002%

Table 20.1 Composition by volume of clean air

Chem-Aid

Noble gases are Group VIII (or Group 0) elements and are also known as inert gases or rare gases. They are a group of monatomic gaseous elements.

The main gases in the air are nitrogen and oxygen. The rest are the noble gases (mostly argon), carbon dioxide and water vapour. The amount of water vapour in air can vary widely around the world, from practically near 0% in a desert to about 5% in a tropical forest.

Air over a busy city also contains toxic gases such as carbon monoxide and sulphur dioxide. Air over a forested area is cleaner.

Fractional Distillation of Liquid Air

Separating air into its constituent gases is an important process, especially for obtaining nitrogen and oxygen. These two gases are widely used in industries. Air is first cooled and compressed into liquid. Liquid air is then separated into its constituents (or fractions) by **fractional distillation** (Fig. 20.1). In fractional distillation, the liquid with the lowest boiling point (b.p.) distils over first. In this case, nitrogen is distilled over first.

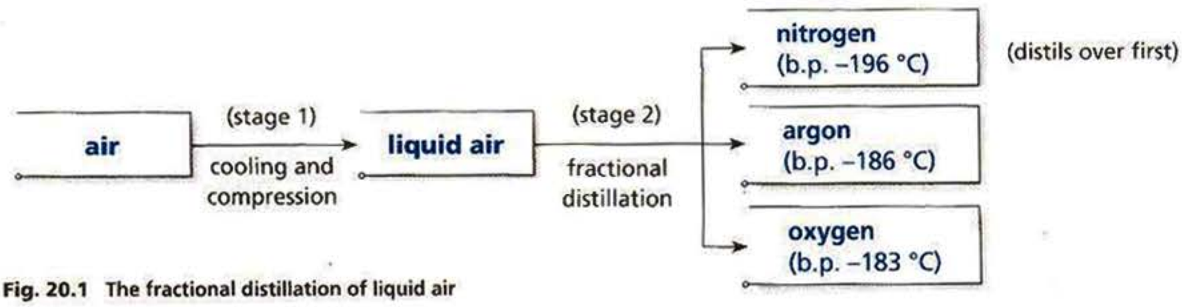


Fig. 20.1 The fractional distillation of liquid air

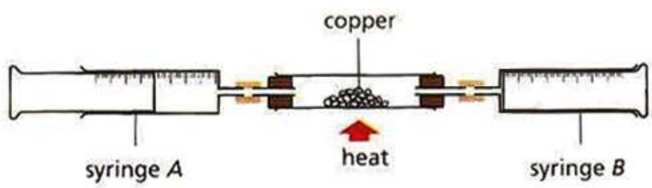
Key ideas

1. Air is a mixture. The composition by volume of clean air is
   - 78 – 79% nitrogen,
   - 20% oxygen,
   - 1% noble gases (mainly argon),
   - 0.03% carbon dioxide,
   - variable amounts of water vapour.
2. Air can be separated into its constituent gases by the fractional distillation of liquid air.

Test Yourself 20.1

Worked Example

200 cm³ of air in syringe *A* was made to pass over heated copper until the reaction was complete.



The gas in syringe *B* was then allowed to cool to its original temperature. What is the volume of gas collected in syringe *B*?

A 80 cm³ C 160 cm³
B 120 cm³ D 400 cm³

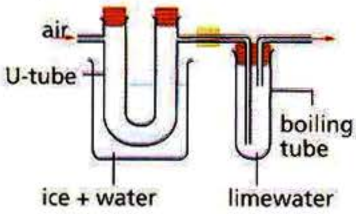
Thought Process

Copper reacted with oxygen in the air to form copper(II) oxide.

copper + oxygen ⟶ copper(II) oxide

$2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$

Since oxygen makes up 20% of air, maximum volume of oxygen that will react = 20% of 200 cm³
= 40 cm³



Therefore, the volume of gas collected in syringe *B*
= 200 cm^3 – 40 cm^3 = 160 cm^3

Answer
C

Questions

1. What will you observe in the (a) U-tube and (b) limewater when air is passed through the apparatus as shown here for a prolonged period? What conclusions about air can you draw from this experiment?
2. In the fractional distillation of liquid air, which gas is distilled over first? Why?

20.2 | Air Pollution

Clean air is essential to life. Unfortunately, much of the air we breathe in is not clean. It contains chemicals that may be harmful to us. **Air pollution** is defined as *the condition in which air contains a high concentration of certain chemicals that may harm living things or damage non-living things.*

Common Air Pollutants

Air pollution is caused by solid particles and poisonous gases in the air. These substances are called **air pollutants**. These pollutants include carbon monoxide, oxides of nitrogen and sulphur dioxide.

Carbon monoxide
Carbon monoxide (CO) is a poisonous gas that is colourless and odourless. It is produced by the incomplete combustion of carbon-containing fuels. Much of the carbon monoxide in the air comes from the incomplete combustion of petrol in car engines.

Oxides of nitrogen
Oxides of nitrogen (NO, NO_2) are produced in two ways:
1. In a car engine or chemical factory, where the temperature is very high, nitrogen combines with oxygen in the air to form nitrogen monoxide (NO). Nitrogen monoxide is also called nitric oxide.

   nitrogen + oxygen ⟶ nitrogen monoxide

   $N_2(g) + O_2(g) \longrightarrow 2NO(g)$

   Nitrogen monoxide, a colourless gas, reacts with oxygen to form a brown gas, nitrogen dioxide (NO_2).

   nitric oxide + oxygen ⟶ nitrogen dioxide

   $2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

2. During thunderstorms, the heat energy released by lightning causes nitrogen and oxygen in the air to react to form oxides such as nitrogen monoxide and nitrogen dioxide.

These oxides dissolve in water to form acid rain. You will learn more about this later.

Chem-Aid

Sulphur dioxide is a reducing agent. In the food industry, it is used as a preservative for foods and drinks.

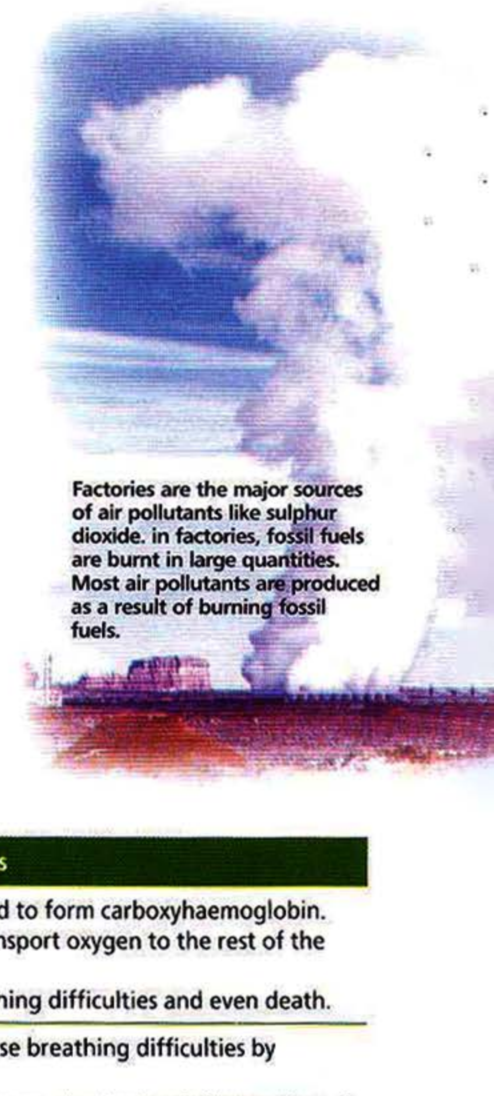
Sulphur dioxide

The major source of sulphur dioxide (SO_2) is the combustion of **fossil fuels** such as coal, crude oil (petroleum) and natural gas. Fossil fuels contain sulphur. When they are burnt, sulphur is converted to sulphur dioxide.

$$\text{sulphur} + \text{oxygen} \xrightarrow{\text{heat}} \text{sulphur dioxide}$$

$$S(s) + O_2(g) \xrightarrow{\text{heat}} SO_2(g)$$

Sulphur dioxide is also produced in large quantities during volcanic eruptions.



Factories are the major sources of air pollutants like sulphur dioxide. in factories, fossil fuels are burnt in large quantities. Most air pollutants are produced as a result of burning fossil fuels.

Other Pollutants

Besides carbon monoxide, oxides of nitrogen and sulphur dioxide, other air pollutants include unburnt hydrocarbons, methane, lead and ozone. Unburnt hydrocarbons and lead are released in car exhaust. Methane is a colourless and odourless gas produced when plant and animal matter decay. It is also produced from the decay of rubbish in landfills.

Methane is a greenhouse gas. You will learn more about greenhouse gases in section 20.4.

Effects of Air Pollution

Air pollution affects human health and the environment in a number of ways. Table 20.2 gives a summary of air pollutants present in the environment and their effects. **Acid rain** in particular is quite harmful.

Air pollutants	Source	Effects
carbon monoxide (CO)	incomplete combustion of fuels such as petrol	• CO reacts with haemoglobin in blood to form carboxyhaemoglobin. As a result, haemoglobin cannot transport oxygen to the rest of the body. • CO causes headaches, fatigue, breathing difficulties and even death.
sulphur dioxide (SO_2)	from volcanoes and combustion of fossil fuels	• These gases irritate the eyes and cause breathing difficulties by irritating the lungs. • High levels of SO_2 and oxides of nitrogen also lead to inflammation of the lungs (bronchitis). • SO_2 and NO_2 form acid rain, which destroys buildings, aquatic life and plants.
oxides of nitrogen (NO, NO_2)	exhaust fumes from vehicles, chemical plants and lightning	

Table 20.2 Effects of air pollutants on human health and the environment

Quick Check

Car exhaust fumes are a major source of air pollution. What pollutants are present in car exhaust?

Acid Rain

Acid rain *is formed when acidic air pollutants such as sulphur dioxide and nitrogen dioxide dissolve in rainwater.* Sulphur dioxide dissolves in water to form **sulphurous acid** (H_2SO_3).

sulphur dioxide + water ⟶ sulphurous acid

$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$

In the presence of oxygen in the air, this acid is slowly oxidised to sulphuric acid (H_2SO_4).

Oxides of nitrogen also contribute to acid rain. In the presence of oxygen and water, nitrogen dioxide is converted to nitric acid.

nitrogen dioxide + water + oxygen ⟶ nitric acid

$4NO_2(g) + 2H_2O(l) + O_2(g) \longrightarrow 4HNO_3(aq)$

The pH value of unpolluted rainwater is usually slightly below 7. This is because carbon dioxide in the air dissloves in rainwater to form carbonic acid, which is a weak acid.

carbon dioxide + water ⟶ carbonic acid

$CO_2(g) + H_2O(l) \longrightarrow H_2CO_3(aq)$

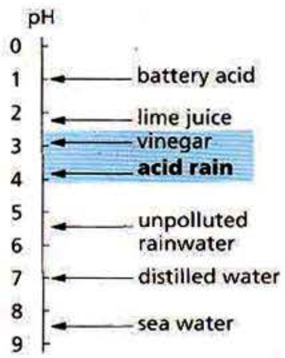


Fig. 20.2 The pH value of acid rain

However, acid rain is much more acidic than rain that only contains carbonic acid. Acid rain has a pH value of 4 or less. Fig. 20.2 shows the pH value of acid rain compared with the pH values of uncontaminated rainwater and other common substances. In some extreme cases, for example in heavy industrial areas, acid rain can be more acidic than vinegar!

What are the effects of acid rain on buildings, plants and aquatic life?

- Acid rain reacts with metals and with carbonates in marble and limestone. When this happens, metal bridges and stone buildings are damaged.
- Acid rain can reduce the pH value of natural water bodies from between 6.5 and 8.5 to below 4, which will kill fish and other aquatic life.
- Acid rain also **leaches** important nutrients from the soil and destroys plants. Without these nutrients, plant growth is stunted. In some cases, acid rain dissolves aluminium hydroxide ($Al(OH)_3$) in the soil to produce Al^{3+} ions, which are toxic to plants. Forests throughout most of Central and Eastern Europe have been destroyed in this manner by acid rain.

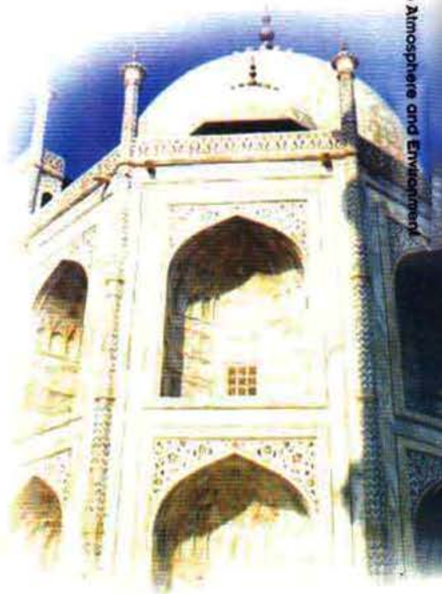
Acid rain causes trees to wither and die.

Key ideas

1. Common air pollutants and their sources:

Air pollutants	Sources
oxides of nitrogen	• lightning activity • inside combustion engines of cars
carbon monoxide	• incomplete combustion of carbon-containing substances, e.g. charcoal, wood, petrol, etc.
sulphur dioxide	• combustion of fossil fuels in motor vehicles, power stations and factories • volcanic eruptions

2. Other air pollutants are methane, lead, ozone and unburnt hydrocarbons.
3. Carbon monoxide can cause breathing difficulties and even death.
4. Nitrogen dioxide and sulphur dioxide in the air react with rainwater to form acid rain.
5. Air pollution can have harmful effects on buildings, human health, plant life and aquatic life.



In India, acidic gas pollutants from nearby industrial plants contribute to the wearing away of the famous marble building, the Taj Mahal.

Test Yourself 20.2

Worked Example

A sample of air in a city was found to contain the following gases: oxides of nitrogen, sulphur dioxide and carbon monoxide.

Which of these gases
a) do not corrode metal structures?
b) are produced during lightning activity?

Thought Process

a) Acidic gases will corrode metal structures. Non-acidic gases will not corrode metal structures. Carbon monoxide is a neutral oxide and is therefore a non-acidic gas.
b) The atmosphere contains nitrogen and oxygen. These two gases will only react at high temperatures to form nitrogen oxides.

Answer

a) Carbon monoxide
b) Oxides of nitrogen

Questions

1. Explain the following observations:
   a) When clean air is bubbled through pure water, the pH of the water changes gradually.
   b) Sitting in a parked car with all the windows closed and the engine running can cause death.

2. a) Identify X in the following equations:

$SO_2(g)$ + water $\longrightarrow$ $X(aq)$

$X(aq)$ + $O_2(g)$ in air $\longrightarrow$ $H_2SO_4(aq)$

b) Balance the following equation:

$SO_2(g) + H_2O(l) + O_2(g) \longrightarrow H_2SO_4(aq)$

c) Explain why acid rain damages buildings and causes farmland to become less fertile.

Depletion of the Ozone Layer

TidBit

The atmosphere is composed of four layers: the troposphere, stratosphere, mesosphere and ionosphere. The air we breathe in is part of the troposphere. Air pollutants are released into this layer. The stratosphere is above the troposphere.

Ozone was discovered in 1840. It was first detected in the upper layer of the atmosphere (the stratosphere) in 1889. It is produced by the photochemical reaction between oxygen molecules and oxygen atoms in the atmosphere.

oxygen atom + oxygen $\longrightarrow$ ozone

$O(g) + O_2(g) \longrightarrow O_3(g)$

When ozone is formed at low altitudes, it can cause severe pollution problems. However, in the stratosphere, ozone becomes very important to us. How is this so?

There is actually a thin layer of ozone in the stratosphere, about 20 – 50 km above the Earth. Here, the ozone acts as a kind of shield, or a giant sunscreen, filtering out some of the harmful ultraviolet radiation from the sun. If this radiation reaches the surface of the Earth, there could be a drastic increase in the number of cases of skin cancer, genetic mutations and eye damage (e.g. cataracts being formed). Ultraviolet radiation may also be harmful to marine life.

Chem-Aid

1. Allotropes are different forms of the same elements.
2. A propellant is a gas that forces out the contents of an aerosol container.
3. A coolant is a liquid that is used for cooling an engine or a mechanical part.

Since 1976, there has been an alarming decrease in the amount of ozone in the stratosphere over the South Pole. In recent years, a similar phenomenon has been occurring over the North Pole.

How is ozone different from oxygen?

Ozone is an allotrope of oxygen. Unlike oxygen, which has two atoms in each molecule, ozone has three atoms of oxygen per molecule. Therefore, its molecular formula is O_3. Ozone is a pale blue, almost colourless gas with a characteristic odour. In small concentrations, it is non-toxic, but in concentrations above 100 ppm (parts per million), it is toxic. Breathing in air that contains high concentrations of ozone can be dangerous, especially for people with asthma.

What risk is this person taking? Are there precautions he can take to minimise the risk to himself?

What is causing the depletion of ozone in the stratosphere?

Scientists have discovered that the depletion of the ozone layer is caused by chlorofluorocarbons. **Chlorofluorocarbons**, commonly called CFCs, are *compounds containing the elements carbon, fluorine and chlorine*. CFCs were widely used as propellants for aerosols and as coolants in refrigerators and air conditioners. They were also used in the manufacture of packing foam.

Within the last few decades, large amounts of CFCs have been released into the atmosphere. CFCs are very stable and can remain in the atmosphere for a very long time. Over the years, they slowly diffuse through the air and react with ozone, destroying the ozone layer.

How do CFCs destroy the ozone layer?

Fig. 20.3 shows how the ozone layer is slowly destroyed by CFCs.

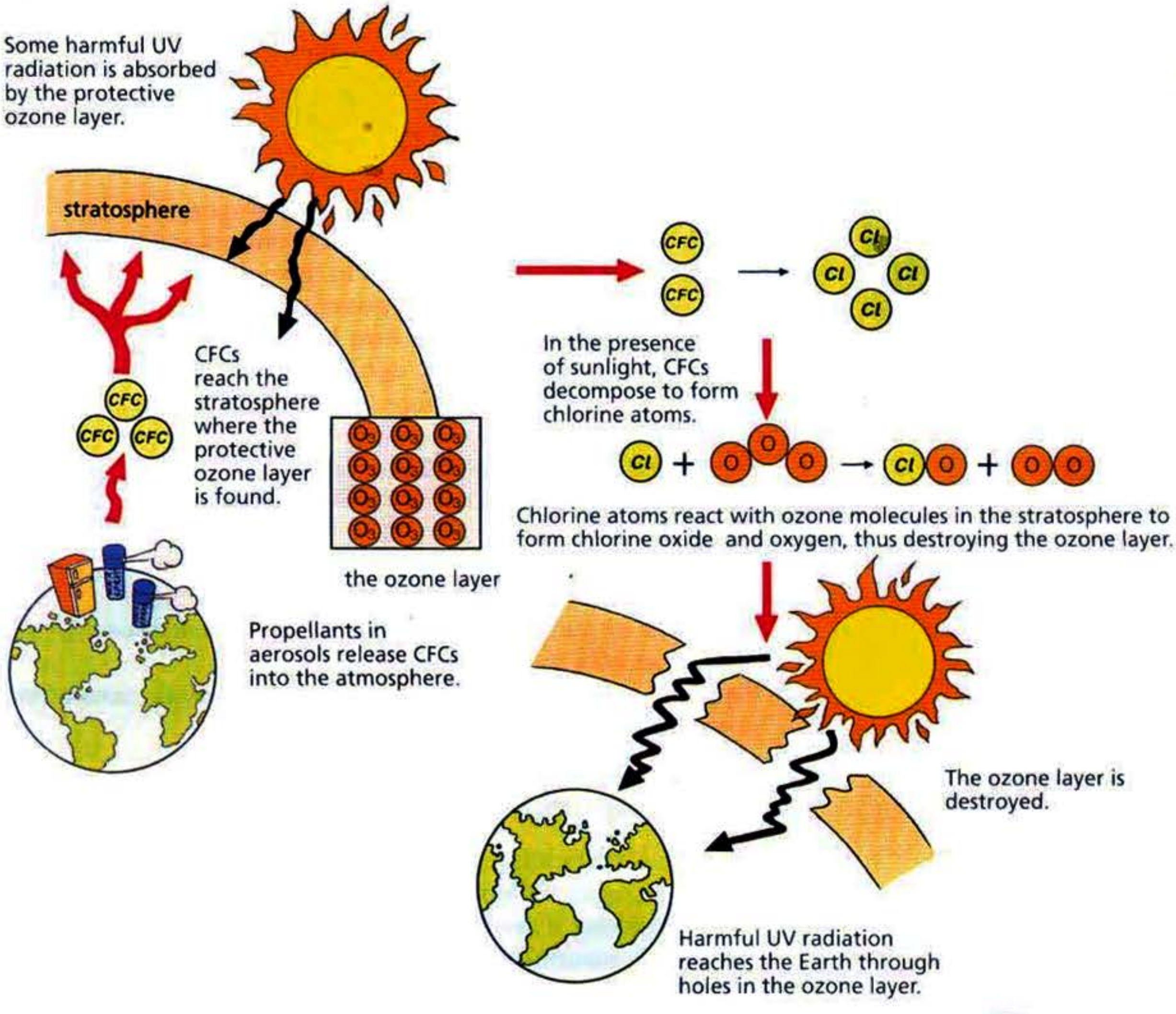


Fig. 20.3 **How the ozone layer is destroyed by CFCs**

Is there a solution to the problem of ozone depletion?

Many countries have now agreed to ban the use of CFCs. In 1992, an international agreement was reached for a complete ban on the release of CFCs by 1996. Until now, most of the countries in the world have completely banned the use of CFCs. However, even if the use of CFCs is totally stopped at once, the depletion of the ozone layer will continue for many years due to the CFCs already present in the atmosphere.

What kind of aerosol sprays do you use at home? Check the product information on the cans. Do these aerosol sprays contain CFCs?

Key ideas

1. The ozone layer filters out some of the harmful ultraviolet radiation from the sun.
2. Chlorofluorocarbons (CFCs) react with ozone in the stratosphere and cause the depletion of the ozone layer.
3. The increase in ultraviolet radiation reaching the Earth could lead to increased cases of skin cancer, genetic mutations and eye damage.

Try it Out

The National Environment Agency (NEA) deals with issues of air and water pollution in Singapore. Visit the NEA's website at www.nea.gov.sg/psi and learn more about
- **how air quality is measured using the PSI index,**
- **the PSI value that represents a dangerous level of air pollution.**

TidBit

Leaded petrol contains small quantities of lead compounds to make the car engines run smoothly. During combustion, lead particles (in the form of volatile lead compounds) are released into the air through the exhaust system, polluting the air. In unleaded petrol, certain hydrocarbons are added, instead of lead, to reduce air pollution.

Chem-Aid

Car tuning is the adjustment made to the fuel supply, air/fuel ratio and spark timing of the car engine. The reason for tuning the engine is to reduce air pollutants in the exhaust gases. An air/fuel ratio of 14 to 15 will give complete combustion of fuels.

20.3 | Reducing Air Pollution

How is air pollution being controlled in Singapore?

The National Environment Agency (NEA) monitors and controls the problems of air and water pollution in the country. The following are some steps taken by the government to control air pollution:

- Prohibition of the use of open fires for the disposal of domestic and industrial wastes. Using open fires to burn domestic and industrial wastes can produce dust, smoke and a significant amount of air pollutants. The air pollutants released depend on the materials being burned. These pollutants are toxic to humans and may cause irritation as well as skin and respiratory problems.
- Introduction of unleaded petrol in 1991, and phasing out of leaded petrol by 1998.
- Reduction of the permissible level of sulphur in diesel (since 1996) from 0.5% to 0.3% by mass.
- Fitting of all petrol-driven vehicles with catalytic converters since 1994.

Reducing the Effects of Acid Rain

The major contributors of acid rain are sulphur dioxide and the oxides of nitrogen. Billions of dollars are lost each year to repair the effects of corrosion due to acid rain. Therefore, it is important to reduce the quantity of these pollutants that is released into the environment. We shall discuss how this is done through the use of catalytic converters and flue gas desulphurisation.

Catalytic Converters

What are the redox reactions in catalytic converters that help to remove pollutants that result from combustion?

Oxides of nitrogen and other undesirable gases (such as carbon monoxide and various unburnt hydrocarbons) emitted by car engines are a major source of air pollution. To reduce air pollution, most cars are now manufactured with catalytic converters.

Fig. 20.4 shows the structure of a catalytic converter. A catalytic converter is attached to the exhaust system of a car. It contains the catalysts platinum and rhodium.

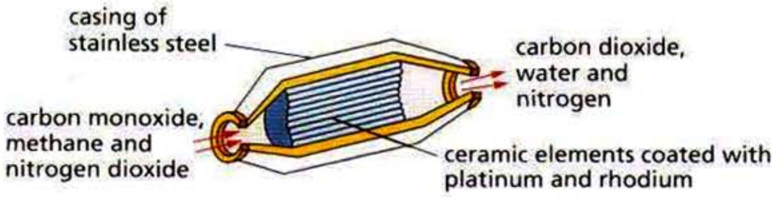


Fig. 20.4 The catalytic converter

When the hot exhaust gases pass over the catalysts, redox reactions occur. The harmful pollutants are converted to harmless substances. For example, carbon monoxide is oxidised to carbon dioxide, oxides of nitrogen are reduced to nitrogen, and unburnt hydrocarbons are oxidised to carbon dioxide and water.

nitric oxide + carbon monoxide ⟶ nitrogen + carbon dioxide

$$2NO(g) + 2CO(g) \longrightarrow N_2(g) + 2CO_2(g)$$

octane + oxygen ⟶ carbon dioxide + water vapour

$$2C_8H_{18}(g) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

What other measures can be used to reduce air pollution caused by motor vehicles?

In the long run, the best way to control air pollution is to burn less fossil fuels. Future control measures may include the use of alternative fuels, such as methanol and hydrogen, and the development of electric-powered motor vehicles.

Flue Gas Desulphurisation

There are two ways of minimising the effects of sulphur dioxide pollution. The most direct approach is to remove sulphur from fossil fuels before they are burnt. However, this method is too expensive and technologically difficult to accomplish. A cheaper way is to remove sulphur dioxide from the waste gases formed when fossil fuels undergo combustion. The waste gases are called **flue gases**. The process of removing sulphur dioxide from flue gases is called **desulphurisation**.

Fig. 20.5 shows a Flue Gas Desulphurisation (FGD) plant. As sulphur dioxide passes through the plant, it reacts with an aqueous suspension of calcium carbonate to form solid calcium sulphite.

calcium carbonate + sulphur dioxide ⟶ calcium sulphite + carbon dioxide

$$CaCO_3(s) + SO_2(g) \longrightarrow CaSO_3(s) + CO_2(g)$$

Quick Check

Suggest three simple ways you can help to reduce air pollution at home. In what ways can chemists help to control air pollution?

Link

Why is methanol called a 'clean' fuel? Find out in Chapter 24.

Quick Check

Instead of removing sulphur dioxide in the form of calcium sulphite, suggest another way of converting sulphur dioxide into a useful starting material for a chemical process.

The calcium sulphite is further oxidised to calcium sulphate by atmospheric oxygen.

calcium sulphite + oxygen → calcium sulphate

$$2CaSO_3(s) + O_2(g) \longrightarrow 2CaSO_4(s)$$

Besides calcium carbonate, calcium oxide can also be used for desulphurisation.

calcium oxide + sulphur dioxide → calcium sulphite

$$CaO(s) + SO_2(g) \longrightarrow CaSO_3(s)$$

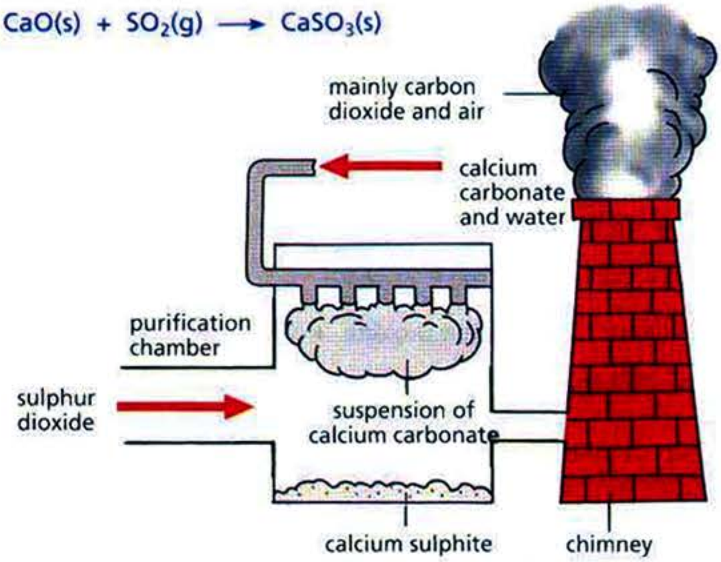


Fig. 20.5 A flue gas desulphurisation plant

Key ideas

1. Catalytic converters are used to convert the oxides of nitrogen, carbon monoxide and unburnt hydrocarbons in car exhaust fumes into harmless substances.
2. Calcium carbonate is used to remove sulphur dioxide from flue gases.

20.4 | The Carbon Cycle

The atmosphere contains 0.03% by volume of carbon dioxide. This represents a huge reservoir of carbon, which is continually being removed from and returned to the atmosphere by a variety of processes.

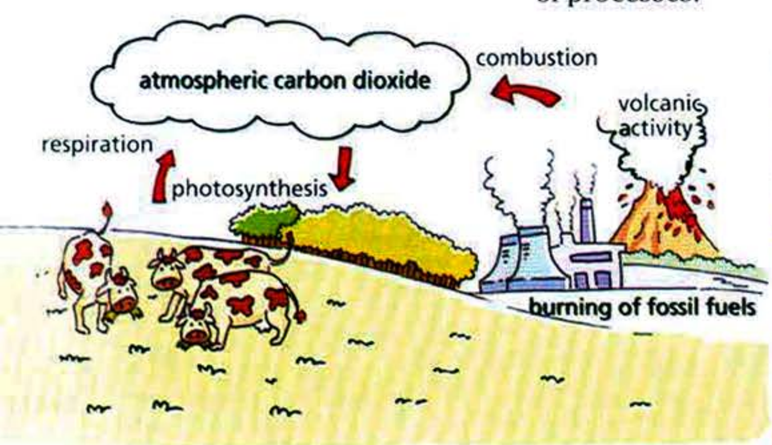


Fig. 20.6 The carbon cycle

If the atmosphere is to maintain a constant amount of carbon dioxide, then the rate of removal of atmospheric carbon dioxide must be balanced by the rate of return of the gas. The mechanism that maintains the level of carbon dioxide in the atmosphere is called the **carbon cycle**.

Fig. 20.6 illustrates the carbon cycle. You can see that there are three processes involved in the cycle: *combustion*, *respiration* and *photosynthesis*.

How is carbon dioxide produced?

There are two main processes which produce carbon dioxide:

1. **Respiration**

   All living things respire. During respiration, carbon and hydrogen compounds in food are converted into carbon dioxide and water. Energy is released during respiration.

$$\text{food} + \text{oxygen} \xrightarrow[\text{cells}]{\text{in living}} \text{carbon dioxide} + \text{water} + \text{energy}$$

$$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l) + \text{energy}$$

2. **Combustion of fuels**

   Coal, petroleum and natural gas are examples of fuels. Most fuels contain both carbon and hydrogen. When these fuels are burnt, carbon dioxide and water are produced. In the combustion of natural gas, which contains mainly methane, the equation for the reaction can be written as

$$\text{methane} + \text{oxygen} \longrightarrow \text{carbon dioxide} + \text{water vapour} + \text{heat energy}$$

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) + \text{heat energy}$$

If a limited supply of air is used, carbon particles (in the form of soot) and the poisonous gas, carbon monoxide, are produced. This is called **incomplete combustion**. The equation for this reaction can be written as

$$\text{methane} + \text{oxygen} \longrightarrow \text{carbon monoxide} + \text{water vapour} + \text{carbon} + \text{heat energy}$$

$$4CH_4(g) + 5O_2(g) \longrightarrow 2CO(g) + 8H_2O(g) + 2C(s) + \text{heat energy}$$

How is carbon dioxide removed from the atmosphere?

Plants are essential to us because they help to remove carbon dioxide from the atmosphere. They do so by the chemical process called **photosynthesis**, which takes place in the green leaves of plants. During photosynthesis, green plants convert carbon dioxide and water into glucose and oxygen (Fig. 20.7) in the presence of sunlight.

Chem-Aid

Photosynthesis is a chemical process by which plants produce food and oxygen. This is an important process because animals cannot make their own food. Humans and other animals depend directly or indirectly on plants for their food and oxygen supply. Chlorophyll, the green pigment in leaves is essential for photosynthesis.

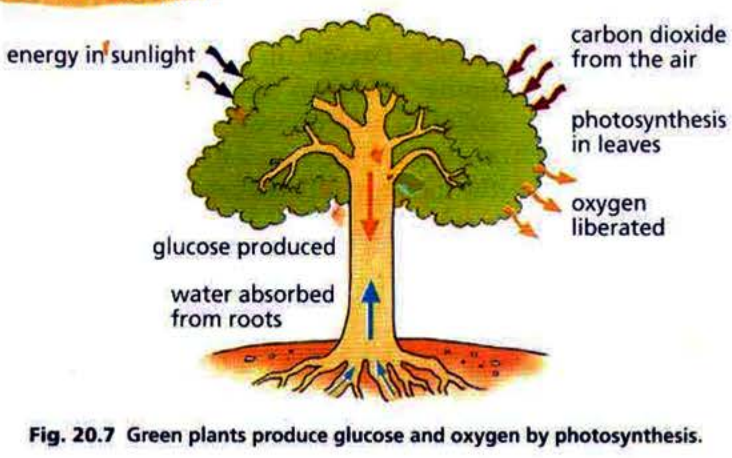


Fig. 20.7 Green plants produce glucose and oxygen by photosynthesis.

Photosynthesis is made up of a series of complex reactions. The overall equation for this process can be written as

carbon dioxide + water $\longrightarrow$ glucose + oxygen

$$6CO_2(g) + 6H_2O(l) \xrightarrow[\text{chlorophyll}]{\text{sunlight}} C_6H_{12}O_6(aq) + 6O_2(g)$$

The glucose formed during photosynthesis is converted by plants into cellulose and starch, which make up the structure and cells of plants.

Photosynthesis is essentially the reverse process of respiration.

glucose + oxygen $\longrightarrow$ carbon dioxide + water

$$\boxed{C_6H_{12}O_6(aq) + 6O_2(g)} \underset{\text{photosynthesis}}{\overset{\text{respiration}}{\rightleftharpoons}} \boxed{6CO_2(g) + 6H_2O(l)}$$

It is important to recognise that the net result of photosynthesis is the conversion of solar energy into a form which can be used by living organisms.

Chem-Aid

Radiation refers to energy travelling in the form of electromagnetic waves or photons.

The Greenhouse Effect

Carbon dioxide and water vapour have important roles to play in maintaining a fairly uniform temperature on the Earth's surface. These gases allow the visible radiation from the sun to reach the Earth's surface but trap some of the infrared radiation which is emitted by the Earth. Energy is thus retained in the atmosphere. This produces a warming effect called the **greenhouse effect**.

The phrase 'greenhouse effect' was first introduced by the French scientist, Jean Fourier, around 1800. He likened the function of the atmosphere to that of the glass in a greenhouse. Sunlight enters the greenhouse and the glass traps some of the sun's radiation energy. The trapped energy helps to keep the greenhouse warm.

Carbon dioxide in the atmosphere has a similar effect as the glass. It is therefore called a greenhouse gas. Carbon dioxide absorbs infrared radiation and thus reduces the amount of heat energy escaping into space. (Fig. 20.8)

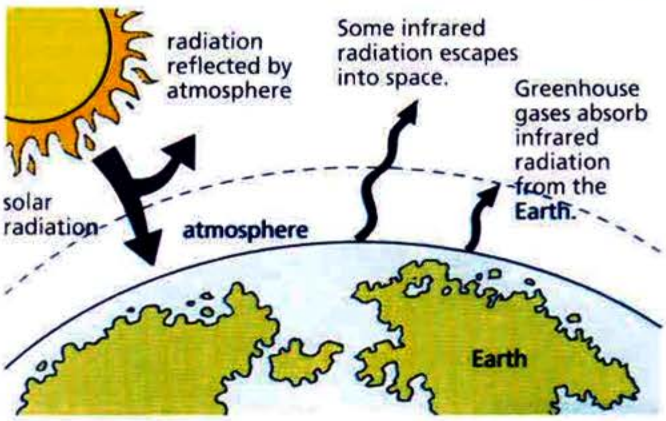


Fig. 20.8 The greenhouse effect

Carbon dioxide is not the only greenhouse gas. Other gases, such as methane (CH_4) and nitrous oxide (N_2O), are even stronger infrared absorbers.

The natural greenhouse effect is crucial for maintaining the proper temperature needed to sustain life on Earth. Without these greenhouse gases, Earth's surface temperature would be -40 °C and it would be permanently covered with ice.

Is the Earth overheating?

Scientists are concerned that the Earth is overheating. Human activities like the burning of fossil fuels and large scale cutting down of forests are causing some greenhouse gases, especially carbon dioxide, to build up rapidly in the atmosphere. This means that carbon dioxide is being added to the atmosphere at a higher rate than photosynthesis can remove the excess gas. The effect of carbon dioxide build-up is an increase in the Earth's average temperature. This phenomenon is called **global warming**.

What are the effects of global warming?

If we do not act to reduce the emission of greenhouse gases, scientists predict that the Earth's temperature could increase by 1 °C to 3 °C within the next 100 years. The possible consequences of global warming include

- a decrease in crop yields world-wide because the areas that are currently covered by vegetation may become deserts,
- the melting of large quantities of ice in the North Pole and South Pole. This will cause the levels of oceans to rise and flood low-lying countries such as the Netherlands.
- the rapid evaporation of water from the Earth's surface. When this happens, carbon dioxide dissolved in the oceans will be driven out into the atmosphere. This adds further to the greenhouse effect.

Quick Check

Suppose the average global temperature was −18 °C instead of a nice comfortable 15 °C. What would human life be like?

Try it Out

What is the Singapore Green Labelling Scheme? How does it contribute in reducing global warming?

Try it Out

Have you heard of the El Nino weather phenomenon? What are the causes and effects of El Nino? You may use the Internet to find out.

Key ideas

1. In the carbon cycle, carbon dioxide is released into the atmosphere by combustion and respiration. Carbon dioxide is removed from the atmosphere by photosynthesis.
2. Carbon dioxide and methane are the main greenhouse gases.
3. Global warming is caused by the build-up of greenhouse gases in the atmosphere.

More and more domestic waste is being buried in landfills. As this waste decays, methane gas, another greenhouse gas, is produced. Suggest one way to reduce the build-up of methane gas in the atmosphere.

Concept Map

Air

amount of CO_2 regulated by

Carbon Cycle

- removal of CO_2 — photosynthesis
- production of CO_2 — combustion and respiration
- rate of CO_2 removal is balanced by rate of CO_2 production

Composition of clean air

- 78–79% nitrogen
- 20% oxygen
- 1% noble gases
- 0.03% carbon dioxide
- variable amount of water vapour

may contain

Air pollutants

- Carbon monoxide
- Nitrogen oxides
- Sulphur dioxide
- Methane

cause

Acid rain

- caused by NO_2 and SO_2
- destroys buildings, plants and aquatic life

Greenhouse effect

- greenhouse gases — CO_2 and methane
- global warming

Ozone depletion

- role of ozone layer — filters out UV light
- ozone depletion — due to destruction of ozone by CFCs

Reducing Air Pollution

Catalytic converter

- catalysts — platinum and rhodium
- NO and NO_2 are reduced to N_2
- CO is oxidised to CO_2
- hydrocarbons are oxidised to CO_2 and H_2O

Flue gas desulphurisation

- SO_2 emitted from burning of fossil fuels
- $CaCO_3$ is used to remove SO_2

Exercise 20

Foundation

1. Which of the following statements about air are true?
   1) Clean air has a higher density than carbon dioxide.
   2) Clean air is a mixture of elements and compounds.
   3) Clean air has a constant composition of oxygen and water vapour.
   4) Clean air contains argon.

   A 1 and 3 only.
   B 2 and 4 only.
   C 1, 2 and 3 only.
   D 1, 2, 3 and 4.

2. The exhaust gases from motor vehicles contain carbon monoxide. Why is carbon monoxide harmful to us?
   A It paralyses the lungs.
   B It corrodes lung tissue.
   C It reduces the ability of blood to carry oxygen to body cells.
   D It makes the blood coagulate.

3. Consider the following gases:
   1) Carbon dioxide
   2) Methane
   3) Nitrogen dioxide
   4) Sulphur dioxide

   Which of these gases may contribute to global warming?

   A 1 only.
   B 1 and 2 only.
   C 2 and 3 only.
   D 3 and 4 only.

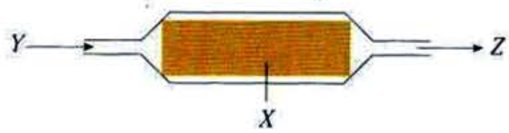
4. a) Define the term 'acid rain'.
   b) i) Name **two** pollutants that cause acid rain.
      ii) What is the similarity between these two pollutants?
      iii) Describe the effects of these air pollutants.
   c) Explain, using appropriate equations, how these substances are involved in the formation of acid rain.

5. a) What is meant by 'desulphurisation'?
   b) Explain briefly the chemistry involved in desulphurisation.
   c) What is the importance of desulphurisation?

Challenge

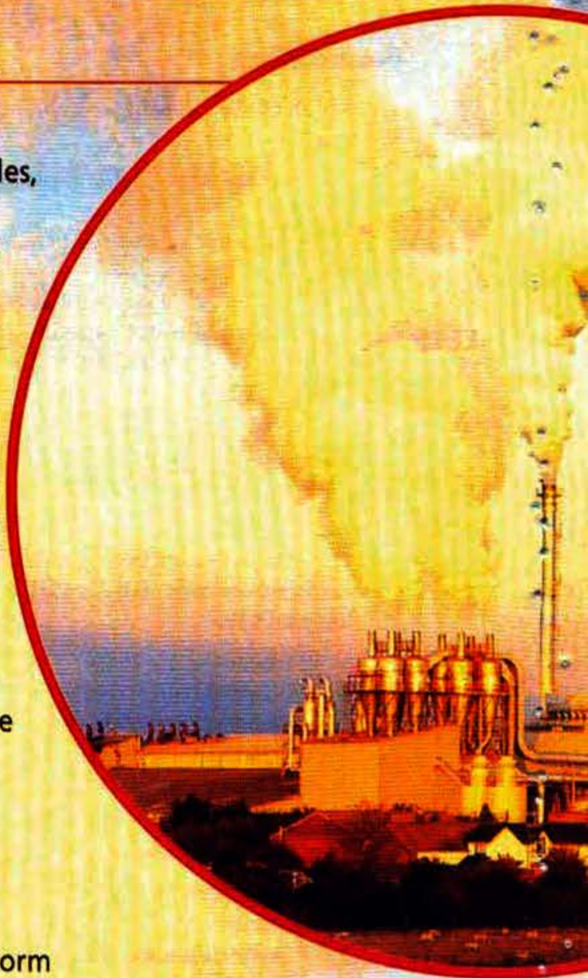
1. A sample of air is bubbled through distilled water. The pH of the solution is 4.8. Which air pollutant causes this?
   A Carbon monoxide
   B Nitrogen dioxide
   C Carbon dioxide
   D Methane

2. The figure below represents a section of a catalytic converter. Catalytic converters are used in car exhaust to reduce air pollution.



   a) i) Name substance *X*.
      ii) What is the purpose of putting *X* in the catalytic converter?
   b) *Y* contains a mixture of gases. Name **three** gases present in *Y* which are
      i) affected by *X*.
      ii) not affected by *X*.
   c) *Z* contains another mixture of gases. Name **three** gases present in *Z* which are produced from chemical reactions as *Y* passes through the catalytic converter.
   d) i) Write the word equations for two reactions that take place in the catalytic converter.
      ii) State the type of reactions that occurs in the catalytic converter.
   e) Suggest why lead-free fuel must be used in cars fitted with a catalytic converter.

Chemistry Today



Photochemical smog is a mixture of pollutants like dust particles, ozone, oxides of nitrogen, unburnt hydrocarbons and peroxyacetyl nitrate (PAN).

How is photochemical smog formed? Nitrogen dioxide in the air reacts with unburnt hydrocarbons in sunlight to produce ozone. Ozone is the main component of photochemical smog. This reaction occurs in two stages. The first stage requires the presence of UV light from the sun to produce oxygen atoms (O).

$$\text{nitrogen dioxide} + \text{oxygen molecules} \xrightarrow{\text{sunlight}} \text{nitrogen monoxide} + \text{oxygen atom} + \text{oxygen molecule}$$

$$NO_2(g) + O_2(g) \xrightarrow{\text{sunlight}} NO(g) + O(g) + O_2(g)$$

The oxygen atoms then react with the oxygen molecules in the air to form ozone (O_3).

$$\text{oxygen atom} + \text{oxygen molecule} \longrightarrow \text{ozone}$$

$$O(g) + O_2(g) \longrightarrow O_3(g)$$

Ozone then further combines with unburnt hydrocarbons to form peroxyacetyl nitrate (PAN).

Photochemical smog can cause headaches and irritation to the eyes, nose and throat. It may also cause coughing and wheezing.

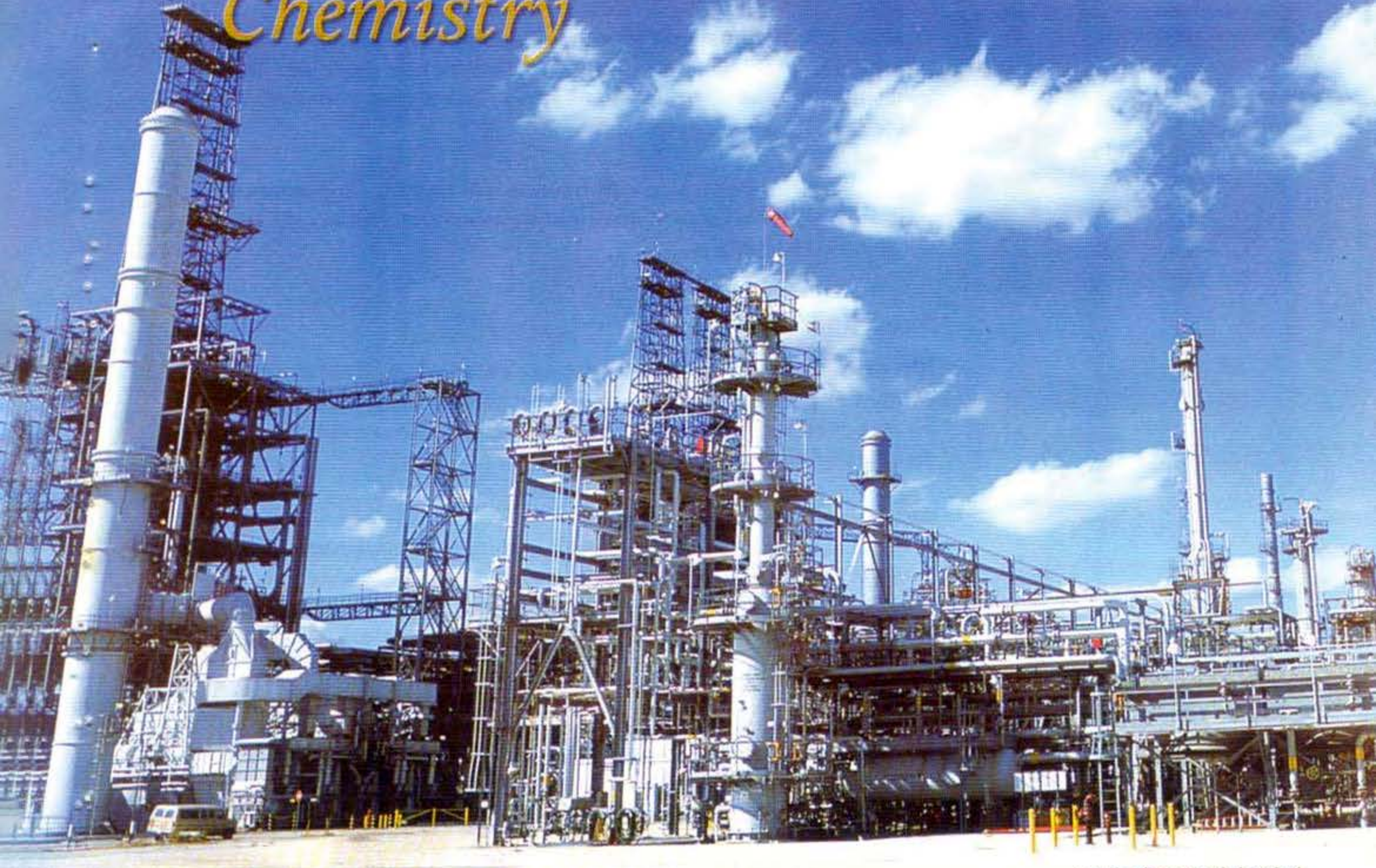
CRITICAL THINKING

Make use of information from the Internet to answer the following:

a) What is the chemical formula of peroxyacetyl nitrate (PAN)?
b) What can we do to reduce the formation of photochemical smog?

Chapter 21

An Introduction to Organic Chemistry



An oil refinery and chemical manufacturing plant.

Chapter Outline

In an oil refinery, petroleum is separated into its components such as kerosene, diesel and petrol. Petroleum is used as a fuel and to manufacture vitamins, plastics, rubber and many other materials important to our modern lifestyle. All these substances contain carbon. In **organic chemistry**, we study substances that contain carbon.

Soaps and detergents are examples of organic compounds.

21.1 | Organic Compounds

What are organic compounds?

All **organic compounds** contain the element carbon. Most organic compounds also contain hydrogen. *Organic compounds that contain only hydrogen and carbon* are called **hydrocarbons**. Organic compounds may also contain other elements such as oxygen and nitrogen.

Organic compounds are found in all animals and plants. Plastics and many medicines are organic compounds because their molecules contain carbon atoms. Even leaves and your hair are made up of organic compounds.

Not all carbon-containing compounds are organic. For example, carbon dioxide, carbon monoxide and carbonates are not classified as organic compounds.

Quick Check

Organic compounds contain carbon and hydrogen. Are they covalent compounds?

Homologous Series

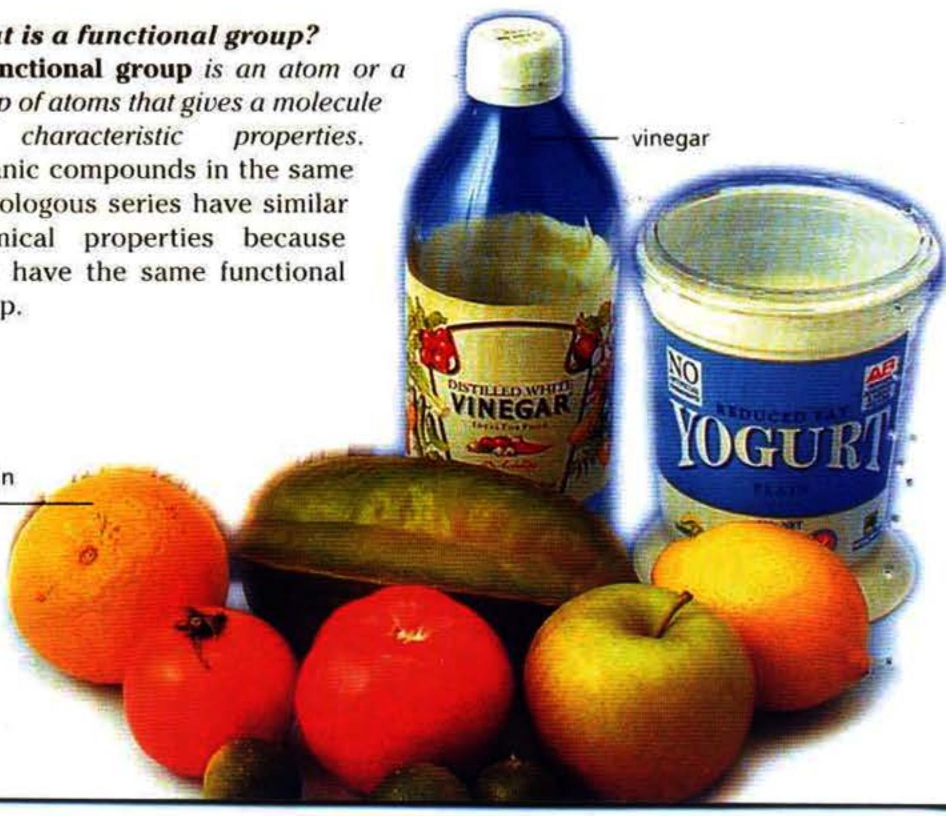
There are millions of different organic compounds. Therefore, the study of organic chemistry would be difficult without some sort of classification. Chemists group related organic compounds into families or **homologous series**.

What is a homologous series?

A homologous series is *a family of organic compounds with similar chemical properties*. Examples of homologous series are the **alkanes**, the **alkenes**, the **alcohols** and the **carboxylic acids**. Compounds of the same homologous series contain the same **functional group**.

What is a functional group?

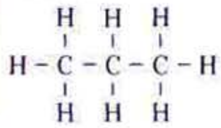
A **functional group** *is an atom or a group of atoms that gives a molecule its characteristic properties*. Organic compounds in the same homologous series have similar chemical properties because they have the same functional group.

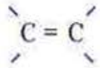
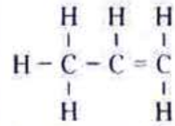


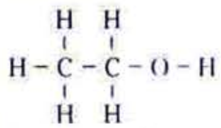
Vinegar and citric acid are carboxylic acids. They belong to the same homologous series and have similar chemical properties.

Table 21.1 shows the four homologous series (alkanes, alkenes, alcohols and carboxylic acids) and their respective functional groups. You will learn more about them in the next few chapters.

Chem-Aid
Alkanes and alkenes are hydrocarbons.

Alkanes	
Functional group: Alkanes do not have any functional groups. There are only C – C and C – H bonds.	Example:  Read more about alkanes in chapter 22.

Alkenes	
Functional group:  carbon-carbon double bond	Example:  Read more about alkenes in chapter 23.

Alcohols	
Functional group: $-O-H$ hydroxyl group	Example:  Read more about alcohols in chapter 24.

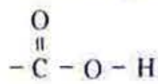
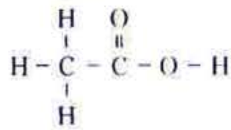
Carboxylic acids	
Functional group:  carboxylic group	Example:  Read more about carboxylic acids in chapter 24.

Table 21.1 Homologous series and their functional groups

What are the general characteristics of a homologous series?
Organic compounds in the same homologous series have the following properties in common:
a) They have the same functional group.
b) They have similar chemical properties.
c) There is a gradual change in their physical properties.

We shall discuss the properties of each homologous series in greater detail in the next few chapters.

Naming Organic Compounds

The name of an organic compound is divided into two parts:
a) The first part (prefix) tells us the number of carbon atoms in each molecule.

First part of the name	meth-	eth-	prop-	but-
Number of carbon atoms in each molecule	one	two	three	four

Table 21.2 Prefixes in naming organic compounds

Alexander Fleming (1881 – 1955)

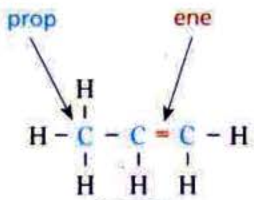
Antibiotics are very important drugs used to treat infections. They are organic compounds. One of the most common antibiotics is penicillin. Penicillin was discovered by Sir Alexander Fleming. During the war between Britain and Germany in 1914, Fleming joined the British Royal Army Medical Corps to develop a cure to reduce the number of soldiers dying from infected wounds. One day in 1928, he made an accidental discovery of a mould growing on the culture of some bacteria. The mould seemed to be able to destroy the bacteria. Fleming went on to show that the fungus contained an antibacterial agent which he called penicillin.

b) The second part (suffix) of the name tells us the homologous series of the compound.

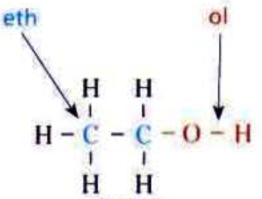
Name of ending	-ane	-ene	-ol	-oic acid
Homologous series	alkane	alkene	alcohol	carboxylic acid

Table 21.3 Suffixes in naming organic compounds

Examples:



propene



ethanol

Thus, propene is an alkene with three carbon atoms per molecule and ethanol is an alcohol with two carbon atoms per molecule.

Key ideas

1. **Organic compounds contain carbon and often contain hydrogen, oxygen and nitrogen as well.**
2. **A hydrocarbon is an organic compound containing only hydrogen and carbon.**
3. **A homologous series is a family of organic compounds with the same functional group and similar properties.**
4. **Examples of homologous series are the alkanes, alkenes, alcohols and carboxylic acids.**
5. **The name of an organic compound is divided into two parts. The prefix indicates the number of carbon atoms in each molecule and the suffix identifies the homologous series.**

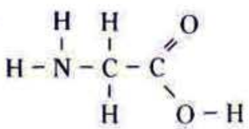
Test Yourself 21.1

Questions

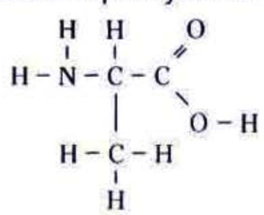
1. Name the functional group in each of the following organic compounds:
   a) Butanol b) Propene c) Ethanoic acid

2. The structures of glycine and alanine are shown below. Do glycine and alanine belong to the same homologous series? Explain your answer.

Glycine



Alanine



21.2 | Petroleum and Natural Gas

Many items that we use in our daily life are actually organic compounds. These include plastics, food, fuels, soaps, disinfectants and even drugs! What is more surprising is that most of these are manufactured from petroleum. In fact, petroleum is a mixture of alkanes.

How are petroleum and natural gas formed?

Millions of years ago, tiny sea creatures and plants sank to the sea bed when they died. The dead creatures and plants were slowly covered by mud and sand. Heat and pressure acted on these organisms over millions of years, eventually producing petroleum and natural gas.

Petroleum and natural gas are often found together in underground deposits hundreds or thousands of metres below the surface of the earth. Deep wells have to be drilled to get them out (Fig. 21.1).

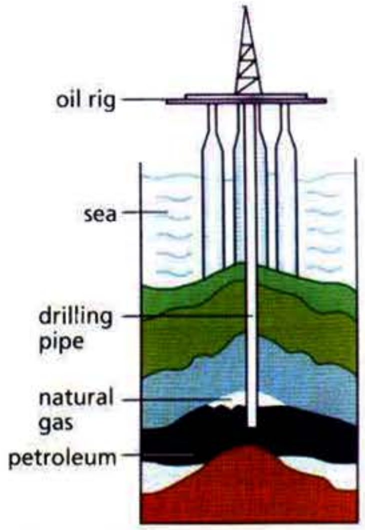


Fig. 21.1 The extraction of petroleum

Fractional Distillation of Petroleum

Petroleum (also called **crude oil**) is a dark brown, foul-smelling liquid. It is a mixture of different hydrocarbons. Petroleum must be separated into fractions (or parts) before it can be useful. *When we* **refine oil**, *we separate petroleum into useful fractions.*

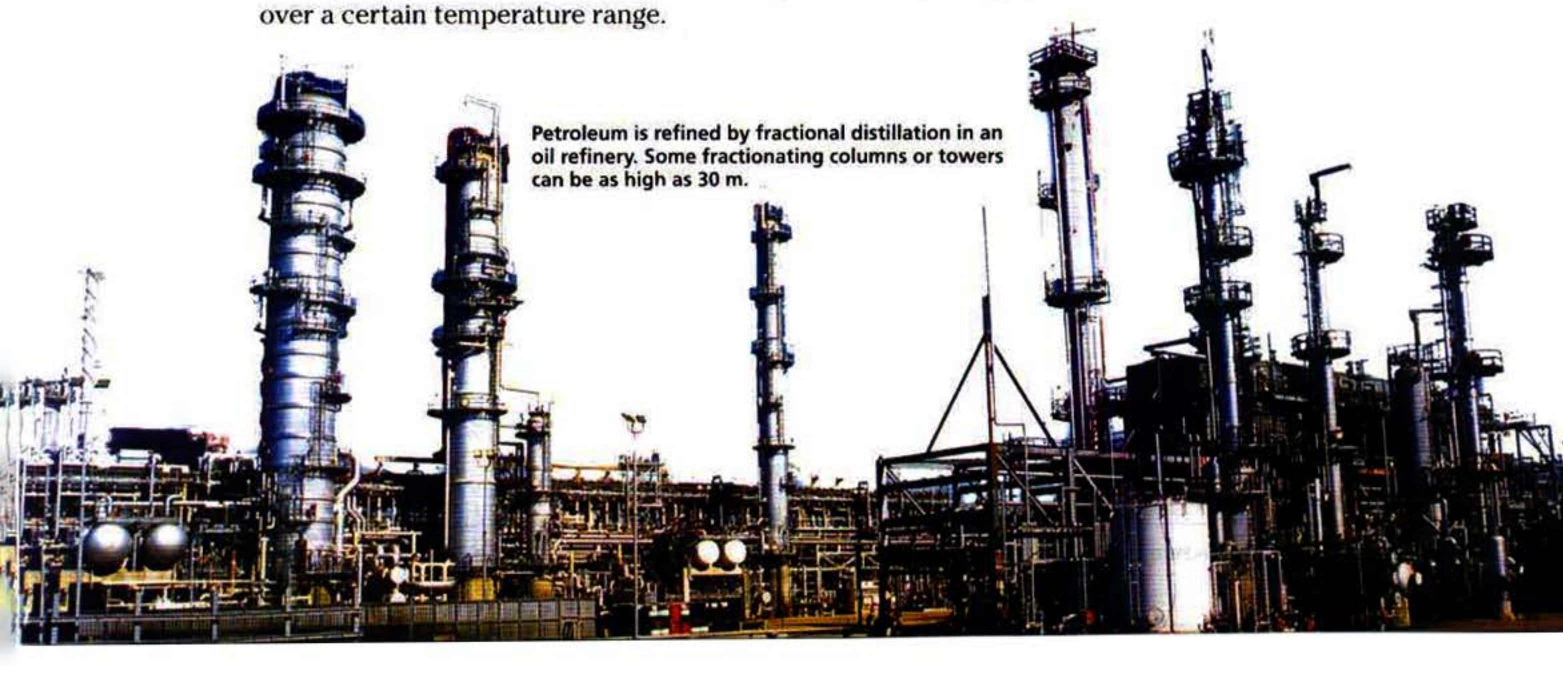
Petroleum can be separated by fractional distillation in an oil refinery (Fig. 21.2). Fractional distillation makes use of the fact that a hydrocarbon with more carbon atoms has a higher boiling point than one with fewer carbon atoms.

Small hydrocarbons (few carbon atoms) have low boiling points and are collected at the top of the fractionating column. The bigger hydrocarbons are collected at the lower sections of the column.

Chem-Aid

Natural gas is mostly made up of methane.

Each petroleum fraction is a mixture of hydrocarbons which boils over a certain temperature range.



Petroleum is refined by fractional distillation in an oil refinery. Some fractionating columns or towers can be as high as 30 m.

What happens in an oil refinery?

Oil refineries separate the hydrocarbons in petroleum by fractional distillation through the following stages:

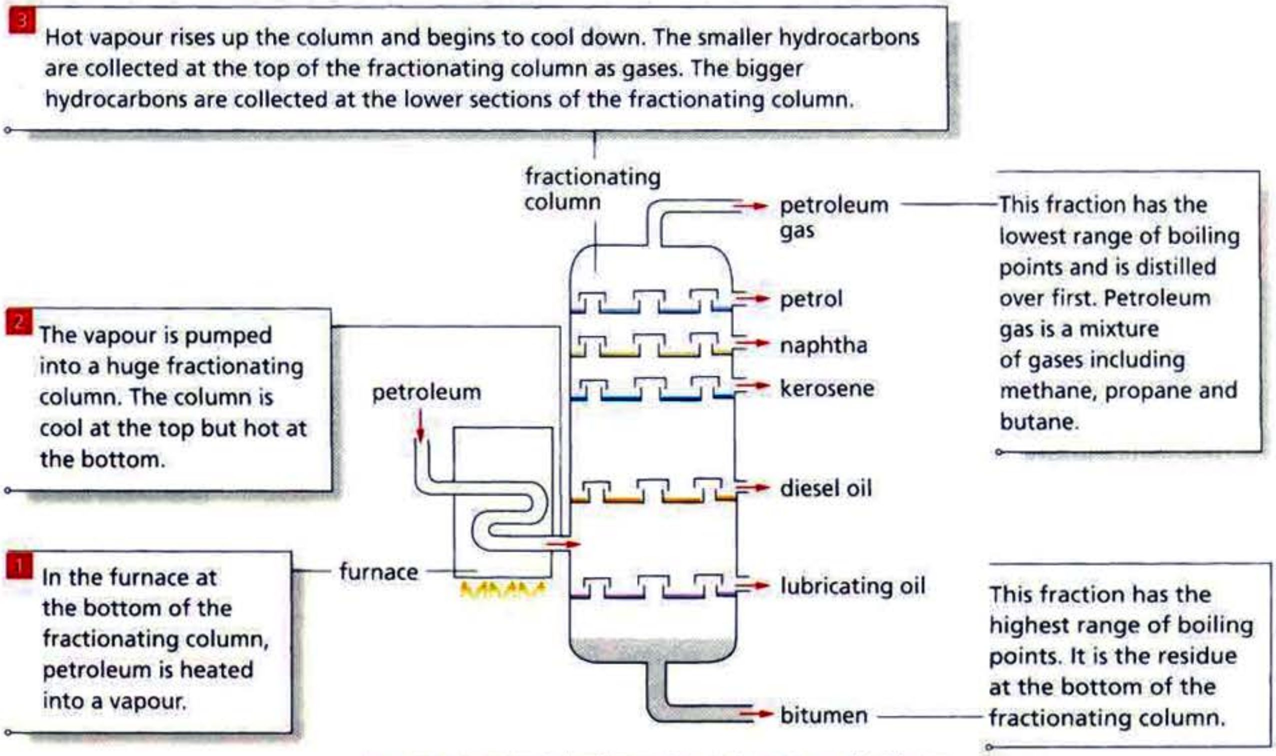


Fig. 21.2 Fractional distillation of petroleum in an oil refinery

The Uses of Petroleum Fractions

The uses of the different petroleum fractions depend on their properties. For example, the petrol fraction burns easily. This makes it a good fuel for petrol engines in vehicles. Some fractions are thick and viscous (do not flow easily), so they are used as lubricating oil for machine joints. Naphtha is used to make other chemicals — we say that naphtha is a feedstock for the petrochemical industry. Table 21.4 summarises the uses of petroleum fractions.

Link

A process called cracking is used to break up larger petroleum molecules into smaller, more useful petroleum molecules. Find out more in chapter 23.

Fractions	Boiling range (°C)	Number of carbon atoms per molecule	Uses
petroleum gas	below 40	1 – 4	fuel for cooking and heating
petrol (gasoline)	40 – 75	5 – 10	fuel for car engines
naphtha	75 – 150	7 – 14	feedstock (raw material) for petrochemical industry (which produces plastics, detergents etc.)
kerosene (paraffin)	160 – 250	11 – 16	fuel for aircraft engines; for cooking using oil stoves; for heating purposes
diesel oil	250 – 300	16 – 20	fuel for diesel engines
lubricating oil	300 – 350	20 – 35	for lubricating machines; for making waxes and polishes
bitumen (asphalt/residue)	above 350	more than 70	for paving road surfaces

Table 21.4 The uses of petroleum fractions

Competing Uses of Petroleum

What are the two main uses of petroleum?

Petroleum is one of the main sources of energy today. About 90% of all the petroleum produced is used as fuel to generate electricity, drive motor vehicles and provide power for industrial activities. Another 10% is used as chemical feedstock (starting materials) for the manufacture of petrochemicals.

Naphtha makes up 20% of petroleum. The naphtha fraction is the main source of hydrocarbons used as the chemical feedstock for the production of a wide range of organic compounds such as plastics, detergents, medicines, synthetic rubber and other useful materials (Fig. 21.3).

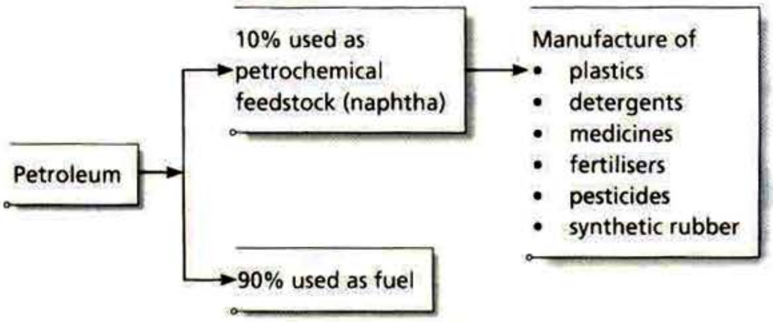


Fig. 21.3 The uses of petroleum in the petrochemical industry

What are the issues relating to the uses of petroleum?

Petroleum is a non-renewable resource and the world's petroleum reserves are finite. With the supply of petroleum decreasing rapidly, there is a growing need for its conservation.

At present, over 90% of the petroleum pumped up from the ground is burnt as fuel. However, petroleum is also important as a chemical feedstock for the manufacture of chemical compounds which are essential for our comfort and health, such as medicines. Therefore, it is a waste to just 'burn' petroleum away. Moreover, burning of petroleum can also cause pollution (see Table 21.6) and global warming (see chapter 20).

The only way to enable petroleum to be used both as a fuel and chemical feedstock is to conserve it. We can cut down on the use of petroleum by reducing the number of motor vehicles on the road, driving smaller cars that consume less petrol, and taking public transport such as buses and the MRT.

We can also save petroleum by using alternative energy sources. Some examples are solar energy and nuclear energy. To produce nuclear energy, nuclear plants use fuels such as plutonium and uranium. However, there is a danger of leakage of harmful radiation into the environment at nuclear plants. Many power stations can also be better designed to use petroleum more efficiently.

TidBit

Fuels can be classified according to their physical states.
- Solid fuels — coal, charcoal, coke
- Liquid fuels — petrol, diesel
- Gaseous fuels — natural gas, hydrogen

Fuels

You have learnt in chapter 17 that a fuel is a substance that can be easily burnt in air to give out energy. We require fuels for activities such as driving motor vehicles and generating electricity.

Coal, petroleum and natural gas are our main sources of fuel. In some countries, wood is also used as a fuel. Most fuels are hydrocarbons.

The general equation for the combustion of any hydrocarbon is shown below.

hydrocarbon (natural gas or petroleum) + oxygen (from air) ⟶ carbon dioxide + water + heat energy

Table 21.5 shows the appearance and constituents of some common fuels.

Fuel	Appearance	Constituents
natural gas	colourless	mainly methane (CH_4)
petroleum (crude oil)	dark brown liquid	a mixture of hydrocarbons, mainly alkanes
coal	black solid	mainly carbon

Table 21.5 Some common fuels

Different fuels have different advantages and disadvantages. When comparing fuels, four important properties are considered: safety, convenience, pollution and amount of energy produced.

The properties of a good fuel are as follows:
a) It must be easy to ignite but not easily flammable.
b) It must be easy to store and transport.
c) It must produce as little pollution as possible.
d) It must produce a lot of heat energy during combustion.

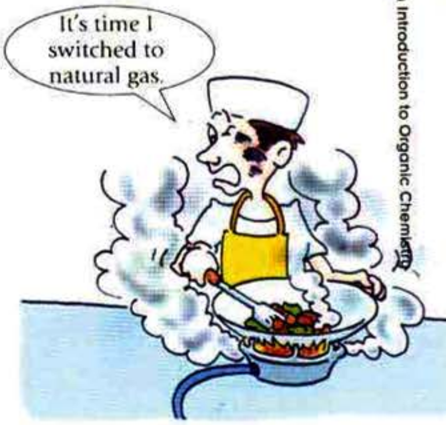
Table 21.6 shows a comparison of coal, natural gas and petroleum as fuels. Which is the best fuel and which is the worst fuel? Explain your choices using the information given in Table 21.6.

Properties	Coal	Petroleum	Natural gas
Physical state	Solid	Liquid	Gas
Ease of storage and transport	Not easily stored or transported	Can be stored and transported easily by tankers Can flow through a pipe easily	Can be liquefied easily and stored in tanks Can flow through a pipe easily
Polluting effect	Extremely polluting. Large amounts of soot and sulphur dioxide produced.	Moderately polluting. Some sulphur dioxide and soot produced.	Almost no pollutants. No soot or sulphur dioxide produced.

Table 21.6 The properties of coal, petroleum and natural gas

Alternative Fuels

The supply of natural gas, petroleum and coal is limited. They are non-renewable resources. These fossil fuels took millions of years to form and we are using them up at a rapid rate.



At this rate of consumption, natural gas and petroleum will run out in less than 50 years. We must conserve these fuels. To do so, we should use these fuels carefully and consider using alternative sources of energy.

One possible source of fuel comes from plants. In Malaysia, experiments using **palm oil** as a substitute for petrol have proven to be successful. Some vehicles fitted with special engines can run on palm oil.

In Brazil, **alcohol** mixed with petrol is used to run vehicle engines. The Brazilians produce this alcohol from sugarcane.

Another important fuel is **biogas**. It is *the gas produced when organic matter (waste material from plants or animals) is allowed to decay in the absence of air.* Biogas contains about 50% methane.

Try it Out

Use the Internet to find out more about
- renewable and non-renewable resources,
- the oil-refining industry in Singapore, and
- how Singapore obtains natural gas from her neighbours.

The Oil Refining Industry in Singapore

Although Singapore is a country with limited natural resources, it is the world's third largest oil-refining centre and a major refining centre for Southeast Asia. The refineries here refine nearly twice the amount of petroleum consumed in Singapore. The rest of the petroleum is exported to different parts of the world.

There are also two industrial plants on Jurong Island, which produce petrochemicals from naphtha fractions.

Key ideas

1. Petroleum (crude oil) and natural gas are important energy sources.
2. Petroleum is a mixture of hydrocarbons.
3. Methane is the main constituent of natural gas.
4. In an oil refinery, petroleum is separated into useful fractions by fractional distillation. The fractions are petrol, naphtha, kerosene, diesel oil, lubricating oil and bitumen.
5. Naphtha is used as a chemical feedstock for the manufacture of organic compounds.
6. Petroleum is a non-renewable resource. It must be conserved if we want to continue using it as a fuel and to manufacture useful products.
7. One way to conserve petroleum is to use alternative sources of fuel such as oils from plants and biogas.

Oil obtained from the fruit of the oil palm is an alternative fuel.

Test Yourself 21.2

Worked Example

LPG (liquefied petroleum gas) is used as a fuel for cooking in homes. Why is LPG stored as a liquid?

Answer

LPG is stored as a liquid so that more fuel can be kept in a container. Liquids contain more particles per unit volume than gases. It is also easier to transport liquids than gases.

Questions

1. Name the method used to refine petroleum. Explain how the method works.

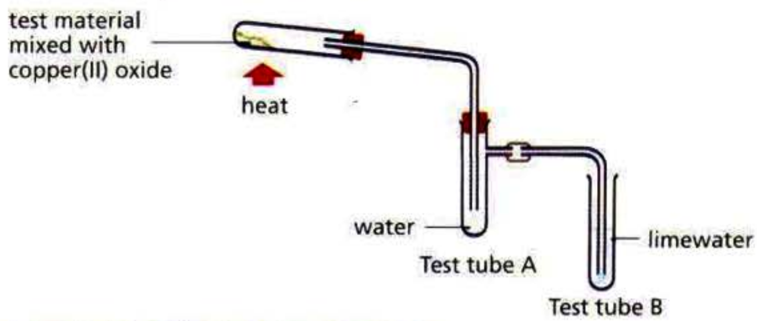
2. Name the oil fraction where the compounds have
   a) the lowest boiling points.
   b) the longest chain lengths.
   State the uses of the fractions in (a) and (b).

3. Petroleum gas is often stored as a liquid. State an advantage and a disadvantage of this method of storage.

4. How does the use of aluminium in the construction of cars help to save fuel?

Science Skills

All organic compounds contain carbon. Almost all organic compounds also contain hydrogen. Hence, the simple apparatus here can be used to show whether a given substance is an organic compound.



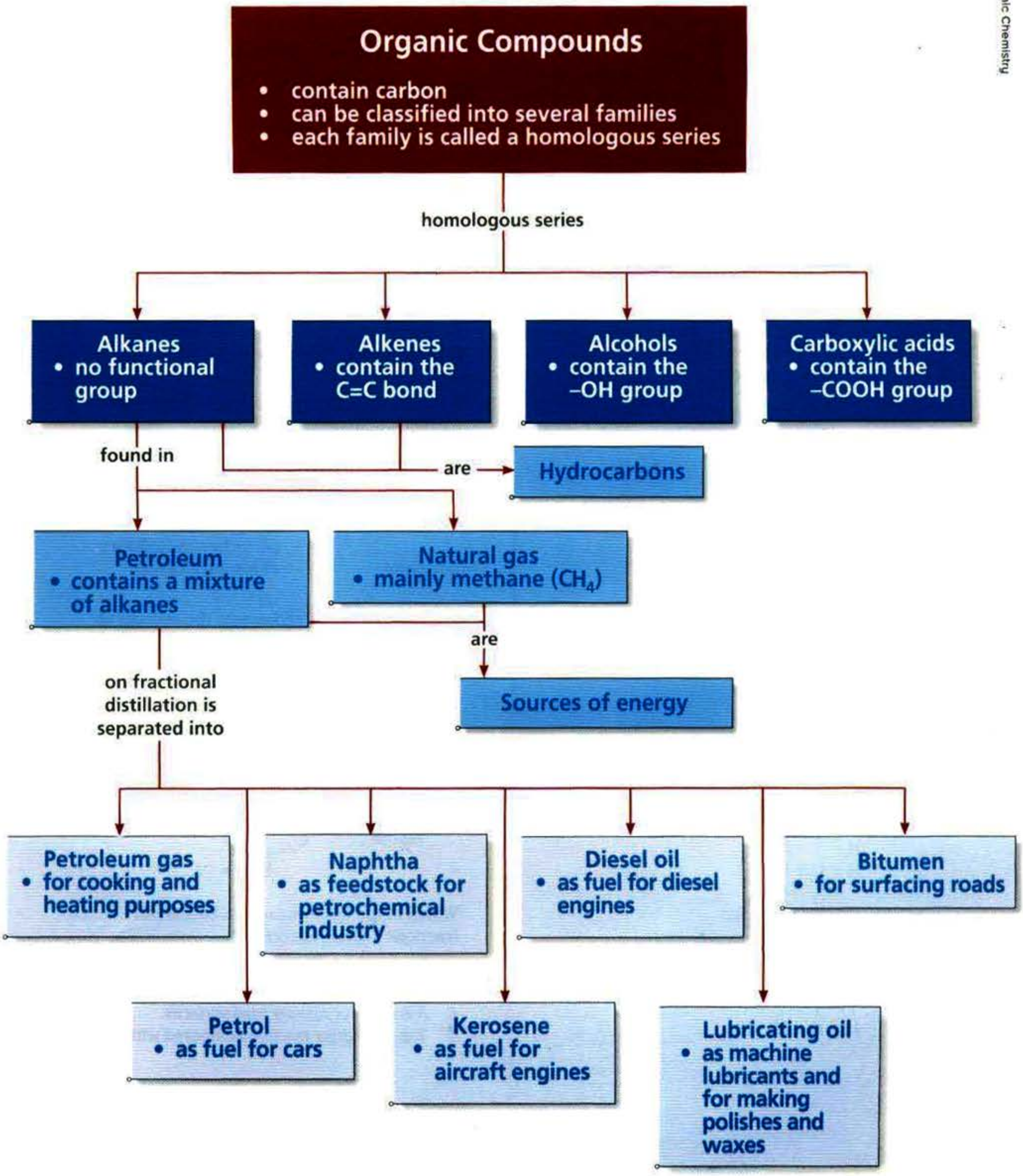
You can use the following as test materials:

sawdust, sugar, fats, starch, wool, cotton, plastics

Heat the substance until the copper(II) oxide turns brown or until there is no further change.

a) What do you observe in test tube B?
b) What is the purpose of water in test tube A?
c) In terms of oxidation and reduction, what happens to copper(II) oxide when it is heated with organic compounds?
d) Most organic compounds contain hydrogen. What will happen to hydrogen in this experiment?

Concept Map


Organic Compounds
• contain carbon
• can be classified into several families
• each family is called a homologous series
homologous series
Alkanes
• no functional group
Alkenes
• contain the C=C bond
Alcohols
• contain the –OH group
Carboxylic acids
• contain the –COOH group
found in
are
Hydrocarbons
Petroleum
• contains a mixture of alkanes
Natural gas
• mainly methane (CH_4)
are
Sources of energy
on fractional distillation is separated into
Petroleum gas
• for cooking and heating purposes
Naphtha
• as feedstock for petrochemical industry
Diesel oil
• as fuel for diesel engines
Bitumen
• for surfacing roads
Petrol
• as fuel for cars
Kerosene
• as fuel for aircraft engines
Lubricating oil
• as machine lubricants and for making polishes and waxes

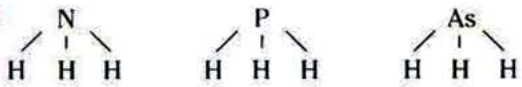
Exercise 21

Foundation

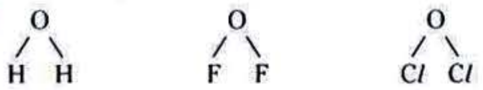
1. Which statement about the members of a homologous series is correct?
   - **A** They have the same solubility in water.
   - **B** They have the same number of carbon atoms per molecule.
   - **C** They undergo similar chemical reactions.
   - **D** They contain the same number of covalent bonds per molecule.

2. In which of the following sets of structures do the members belong to the same homologous series?

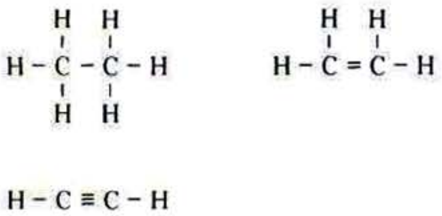
A



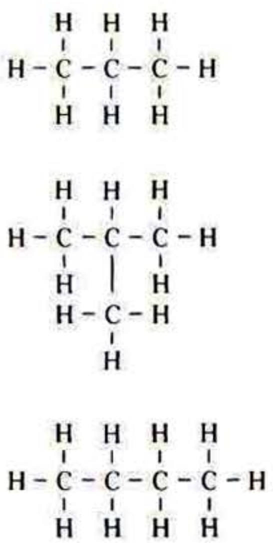
B



C



D



3. Which of the following compounds is named correctly?

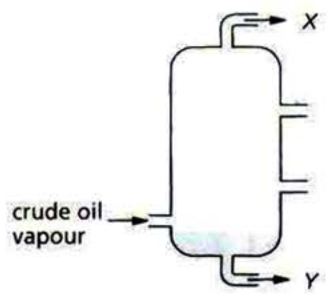
	Formula	Name
A	CH_3COOH	methanoic acid
B	CH_3CH_2COOH	ethanoic acid
C	$CH_3CH_2CH_2OH$	propanol
D	CH_3CH_2CHO	propanol

4. When petroleum is distilled, several fractions can be collected. Which of the following lists arranges the fractions from the lowest to the highest boiling point?
   - **A** Diesel, kerosene, petrol, lubricating oil.
   - **B** Petrol, kerosene, diesel, lubricating oil.
   - **C** Kerosene, petrol, lubricating oil, diesel.
   - **D** Lubricating oil, kerosene, diesel, petrol.

5. Which of the following petroleum fractions is correctly matched with its use?

	Fraction	Use
A	bitumen	for making roads
B	naphtha	for making polishes
C	paraffin	as feedstock for petrochemical industry
D	diesel	as a source of wax

6. The diagram below represents the process of fractional distillation of crude oil.



Which of the following statements about fractions X and Y is correct?
   - **A** X burns more easily than Y.
   - **B** X has a higher boiling point than Y.
   - **C** X is used for making road surfaces.
   - **D** Y is the lighter fraction compared with X.

7. In the fractional distillation of petroleum,
   A the hydrocarbons collected in each fraction have similar boiling points.
   B the fractions with higher boiling points are collected higher up the fractionating column.
   C the fractions are separated according to their melting points.
   D the fractions collected at the top are the alkanes, followed by the alkenes.

8. Which of the following organic compounds would you expect to find in naphtha?

   A C_2H_4 B C_3H_8
   C C_8H_{18} D $C_5H_{12}O$

9. Why is natural gas suitable for use as a fuel?
   A It has a low boiling point.
   B It burns exothermically.
   C It is a gas.
   D It has a high content of carbon.

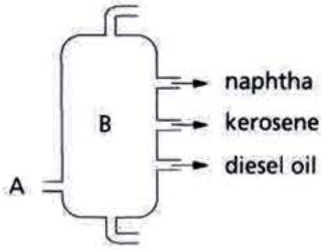
10. State **three** general characteristics of organic compounds that belong to the same homologous series.

11. Consider the following petroleum fractions:

diesel oil, kerosene, naphtha, petroleum gas, bitumen

   a) Place the above fractions in order of decreasing volatility (most volatile to least volatile).
   b) Of the fractions given above, which
      i) contains hydrocarbons with greater than 20 carbon atoms per molecule?
      ii) contains propane (C_3H_8)?
      iii) is used as a starting material for making plastic?

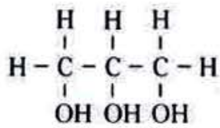
12. The figure below shows how crude oil (petroleum) can be refined.



   a) What does 'refining crude oil' mean?
   b) Name the process used to refine crude oil.
   c) What change of state occurs at 'A'?
   d) Explain how crude oil is separated at 'B'.
   e) State (i) **two** similarities and (ii) **two** differences between naphtha and diesel oil.
   f) Name two fuels, suitable for cars, which are not obtained from crude oil.

Challenge

1. Glycerine is used as an anti-freeze in car radiators during the winter season. The formula of glycerine is



   The chemical properties of glycerine are expected to be similar to that of ________.

   A propane B propene
   C propanol D propanoic acid

2. A hydrocarbon was found to contain 17.2% hydrogen by mass. Its empirical formula is

   A CH_2 B C_2H_2
   C C_2H_4 D C_2H_5

Chapter 22

Alkanes

Machinery on fire

Petrol is used as a fuel to power cars, ships and many other types of machinery. However, petrol is also extremely dangerous. The photograph shows what happens when machinery running on petrol catches fire. Why is petrol able to burn so easily?

The answer lies in the hydrocarbons that make up petrol. In this chapter, we will learn more about the hydrocarbons in petrol — the alkanes.

Chapter Outline

22.1 | Structure of Alkanes

Alkanes are hydrocarbons that contain only single covalent bonds between the carbon atoms. In chapter 21, you learnt that alkanes form a homologous series. Table 22.1 lists the first five members of the alkane homologous series.

Name	Number of carbon atoms	Molecular formula	Melting point (°C)	Boiling point (°C)	Physical state at r.t.p.
methane	1	CH_4	−182	−162	gas
ethane	2	C_2H_6	−183	−89	gas
propane	3	C_3H_8	−188	−42	gas
butane	4	C_4H_{10}	−138	−0.5	gas
pentane	5	C_5H_{12}	−130	36	liquid

Table 22.1 The first five members of the alkane homologous series

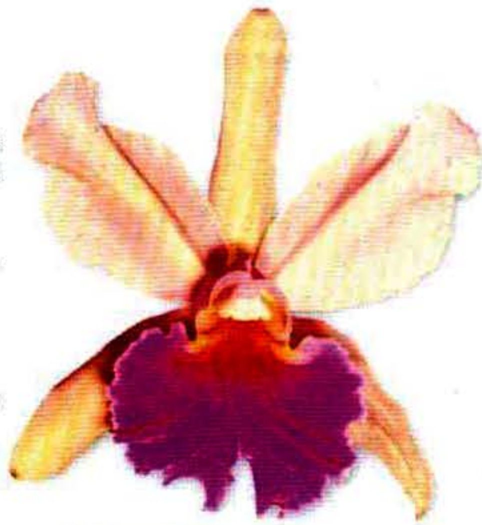
From the table, it is clear that each member of the alkane homologous series differs from the next by a $-CH_2-$ unit. For example, ethane differs from methane by having an additional $-CH_2-$ unit.

$$\underset{\text{methane}}{CH_4} + -CH_2- \longrightarrow \underset{\text{ethane}}{C_2H_6}$$

What is the general molecular formula for the alkane homologous series?

Alkanes have the general molecular formula $\mathbf{C_nH_{2n+2}}$. If an alkane contains one carbon atom, n = 1. It has the molecular formula $C_1H_{(2\times1)+2} = CH_4$. This is methane.

Similarly, the alkane with two carbon atoms has the molecular formula $C_2H_{(2\times2)+2} = C_2H_6$. This is ethane.



Some orchids produce alkanes (e.g. $C_{23}H_{48}$) to attract bees to pollinate their flowers!

What are the properties of a homologous series?

Previously, you learnt that members of the same homologous series have three general properties. They
- have the same functional group,
- have similar chemical properties,
- show a gradual change in their physical properties.

Based on what we now know about alkanes, we can add two more general properties of a homologous series, i.e.
- members of the same homologous series have the same general formula,
- each member of the series differs from the next by a $-CH_2-$ unit.

What is the structural formula of an organic compound?

A **structural formula** is *the formula which shows how atoms are arranged in a molecule.* The **full structural formula** shows all the bonds between the atoms in a molecule. This allows us to see the arrangement of the atoms in a molecule.

Table 22.2 shows the various formulae, and the three-dimensional models of the first three members of the alkane series.

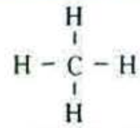
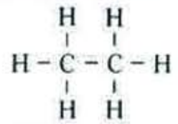
Name	Molecular formula	Structural formula	Full structural formula	Three-dimensional model
Methane	CH_4	CH_4		
Ethane	C_2H_6	CH_3CH_3		
Propane	C_3H_8	$CH_3CH_2CH_3$		

Table 22.2 The formulae and models of some alkanes

The carbon atom in methane forms four single bonds.

Why are alkanes called saturated hydrocarbons?

Alkanes are called **saturated hydrocarbons** because they *contain only single carbon-carbon covalent bonds*. Each carbon atom in an alkane molecule uses all its valence electrons in forming single bonds with four other atoms. For example, the carbon atom in methane has four electrons in its outer shell. It has formed four single bonds with hydrogen, thus attaining the stable electronic configuration of a noble gas.

Fig. 22.1 The electronic structure of methane

What are the electronic structures of methane and ethane?

The electronic structures of methane and ethane are shown in Fig. 22.1 and 22.2 respectively.

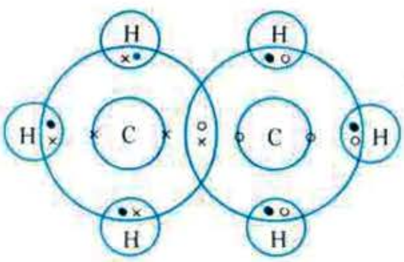
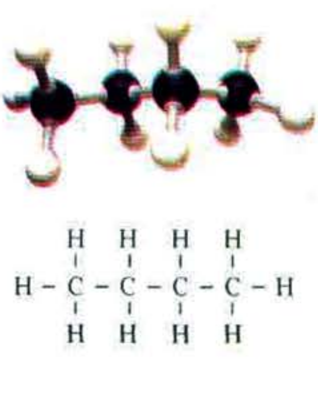


Fig. 22.2 The full electronic structure of ethane

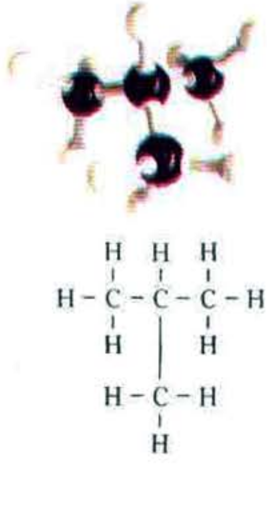
Isomerism in Alkanes

Look carefully at the following hydrocarbons, A and B. How are they similar or different?

A



B

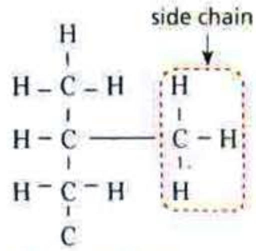


Hydrocarbons A and B have the same molecular formula, C_4H_{10}. We can see that the arrangement of the atoms of A and B are different. Hence, they have different structural formulae. Hydrocarbon A is called butane and B is called methylpropane.

Butane and methylpropane are **isomers**. *Compounds that have the same molecular formula but different structural formulae* are called isomers.

It is important to realise that isomers are different chemical compounds. They have different structures. They also have different physical properties such as melting and boiling points. For example, the boiling point of butane is −0.5 °C while the boiling point of methylpropane is −11.7 °C.

Butane is a **straight-chain** alkane because all the carbon atoms are joined up in a row. Methylpropane is a **branched-chain** alkane because it has a branch or a side chain, CH_3 (Fig. 22.3).



Methylpropane

Fig. 22.3 Methylpropane has a side chain.

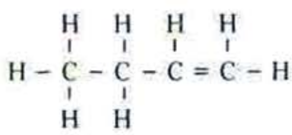
The side chains are called **alkyl groups**. An alkyl group is obtained by removing a hydrogen atom from an alkane. It therefore has the general formula $\mathbf{C_nH_{2n+1}}$. Typical alkyl groups are

- methyl: CH_3-
- ethyl: CH_3CH_2-
- propyl: $CH_3CH_2CH_2-$
- butyl: $CH_3CH_2CH_2CH_2-$

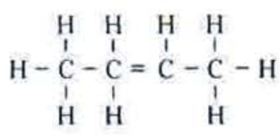
Test Yourself 22.1

Worked Example

Two hydrocarbons, *A* and *B*, are shown below. Are *A* and *B* isomers of each other?



A



B

Thought Process

The molecular formulae of *A* and *B* are the same (C_4H_8). Their structural formulae are different as the functional groups are at different locations.

Answer

Yes. *A* and *B* are isomers of each other. (Both are butene.)

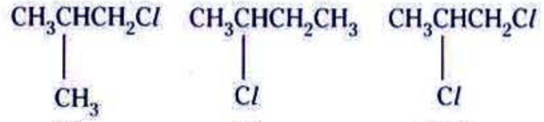
Questions

1. The hydrocarbon, dodecane, is used to make detergents. It is an alkane and has the molecular formula C_xH_{26}. What is the value of x?

Key ideas

1. Alkanes are hydrocarbons with the general formula C_nH_{2n+2}.
2. Alkanes contain only single carbon-carbon covalent bonds. They are known as saturated hydrocarbons.
3. Each member of the alkane homologous series differs from the next by a $-CH_2-$ unit.
4. Isomers are compounds that have the same molecular formula but different structural formulae.

2. Select from the list below the isomer(s) of $CH_3CH_2CH_2CH_2Cl$.



(I) (II) (III)

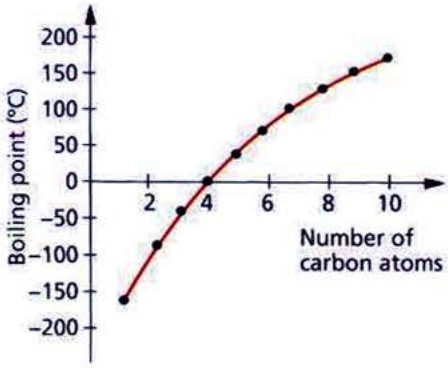


Fig. 22.4 The boiling points of some alkanes

22.2 | Properties of Alkanes

All alkanes are insoluble in water. They all burn in air (are flammable). Alkanes also show a gradual trend in the following properties: melting point and boiling point, density, viscosity and flammability.

Melting Point and Boiling Point

The melting and boiling points of the alkanes increase as their molecular sizes increase. This is because as alkane molecules become bigger, the attractive forces between the alkane molecules become stronger.

Density

The density of alkanes increase as their molecular sizes increase.

Viscosity

The alkanes become more *viscous*, i.e. more difficult to pour out, as their molecular sizes increase. This is because the intermolecular forces are stronger as the alkane molecules become bigger. This makes it difficult for the liquid to flow.

Flammability

As the molecular size of the alkane molecules increases, the percentage of carbon in the alkane molecules also increases. As a result, alkanes become *less flammable*, i.e. more difficult to burn. They also produce a smokier flame. The smoky flame is caused by the incomplete combustion of alkane molecules.

Combustion

The most important property of alkanes is that they burn readily in air when ignited by a spark or flame. Combustion of alkanes *in excess air or oxygen* produces carbon dioxide and water vapour.

TidBit

The waxy coating on fruits helps to retain moisture and prevent fruits from spoiling. This waxy coating is made of solid alkanes. Estimate the number of carbon atoms in these solid alkanes.

The equation for the complete combustion of methane is

$$\text{methane} + \text{oxygen} \xrightarrow{\text{heat}} \text{carbon dioxide} + \text{water vapour}$$

$$CH_4(g) + 2O_2(g) \xrightarrow{\text{heat}} CO_2(g) + 2H_2O(g)$$

If there is an *insufficient supply of air or oxygen*, combustion of alkanes is incomplete. As a result, carbon and carbon monoxide may be formed too.

$$\text{methane} + \text{oxygen} \xrightarrow{\text{heat}} \text{carbon monoxide} + \text{water vapour}$$

$$2CH_4(g) + 3O_2(g) \xrightarrow{\text{heat}} 2CO(g) + 4H_2O(g)$$

These reactions are highly exothermic, hence alkanes make good fuels. For example, methane is used in cooking gas. It is colourless and odourless. Traces of a foul smelling chemical are usually added to cooking gas so that it can be detected. Other examples of alkanes which are used as fuels are petroleum gas, kerosene and diesel oil. The bigger alkane molecules, for example, those that make up candle wax, also burn. However, they burn with a sooty flame.

Lighters contain alkane molecules. These molecules exist as a liquid in such containers.

Quick Check

Write down the equation for the complete combustion of ethane.

Cracking

Alkanes with long carbon chains such as those found in bitumen, are not usually used as fuels. This is because they are less flammable. However, alkanes with long carbon chains can be broken up into smaller molecules. This process is called **cracking**. You will learn more about cracking in chapter 23.

Substitution Reactions

Why are alkanes generally unreactive?

Alkanes are saturated hydrocarbons with single carbon-carbon (C–C) bonds and carbon-hydrogen (C–H) bonds. These bonds are strong and are difficult to break. As a result, alkanes are generally unreactive.

However, alkanes do react with halogens such as chlorine and bromine in the presence of UV (ultraviolet) light. These are called **substitution reactions**.

Methane reacts with chlorine in the presence of UV light as follows:

$$\text{methane} + \text{chlorine} \xrightarrow{\text{UV light}} \text{chloromethane} + \text{hydrogen chloride}$$

$$CH_4(g) + Cl_2(g) \longrightarrow CH_3Cl(g) + HCl(g)$$

$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array} + Cl-Cl \longrightarrow \begin{array}{c} H \\ | \\ H-C-Cl \\ | \\ H \end{array} + H-Cl$$

In this reaction, a hydrogen atom in methane is replaced (substituted) by a chlorine atom to form chloromethane. Hence the reaction is called a **substitution** reaction. Hydrogen chloride is also produced.

More hydrogen atoms can be replaced with chlorine atoms to form other compounds such as dichloromethane (Fig. 22.5)

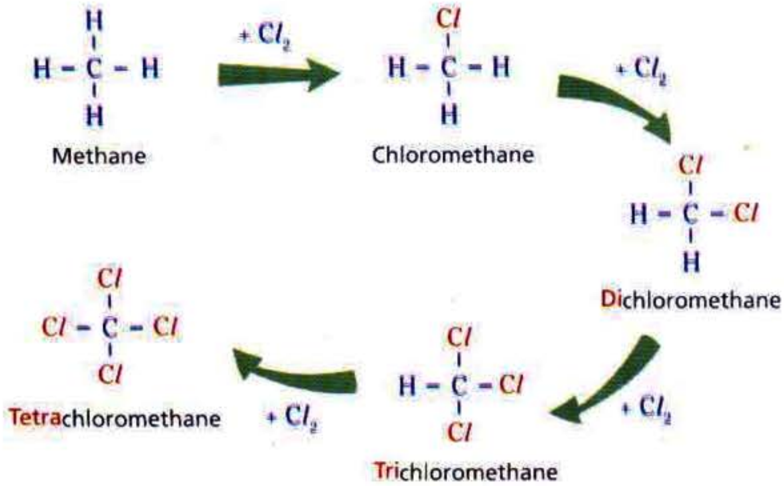


Fig. 22.5 The substitution reaction of methane and chlorine produces four products.

Key ideas

1. As the molecular sizes of the alkanes increase,
   - their densities, melting and boiling points increase,
   - they become more viscous,
   - they become less flammable.
2. Alkanes are generally unreactive.
3. Alkanes burn in excess oxygen to form carbon dioxide and water vapour. When there is insufficient oxygen, carbon and carbon monoxide are also produced.
4. Alkanes with long carbon chains can be broken up into smaller molecules by cracking.
5. Alkanes undergo substitution reactions with chlorine in the presence of UV light to give a mixture of products.

Test Yourself 22.2

Worked Example

You are given two liquid hydrocarbons *A* and *B*.
Hydrocarbon *A*: molecular formula C_6H_{14}
Hydrocarbon *B*: molecular formula $C_{16}H_{34}$

Two drops of each hydrocarbon are ignited. Which of these two hydrocarbons gives out more smoke?

Thought Process

The more carbon atoms there are in a hydrocarbon molecule, the more smoke will be given off as the hydrocarbon burns.

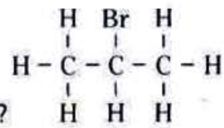
Answer

Hydrocarbon *B* gives out more smoke.

Questions

1. On complete combustion, one mole of an alkane produces two moles of carbon dioxide and three moles of water. Suggest the molecular formula of this alkane.

2. A hydrocarbon *Y* reacts with bromine to form a compound *Z* and hydrogen bromide. The full structural formula of *Z* is shown here. What is the name and full structural formula of *Y*?



3. A hydrocarbon X has the molecular formula C_8H_{18}. What is the molecular formula of the hydrocarbon Y which has two carbon atoms less than X in the same homologous series? Give **two** properties in which (a) X is similar to Y, and (b) X differs from Y.

4. The substances X, Y and Z shown in the following reactions are alkanes.

$$X + \text{chlorine} \xrightarrow{\text{UV light}} CH_3CH_2Cl + HCl$$

$$Y + \text{chlorine} \xrightarrow{\text{UV light}} CH_3\underset{\underset{\displaystyle CH_3}{|}}{C}HCH_2Cl + HCl$$

$$Z + \text{chlorine} \xrightarrow{\text{UV light}} CH_3\underset{\underset{\displaystyle Cl}{|}}{C}HCH_2CH_3 + HCl$$

Identify and draw the full structural formulae of X, Y and Z.

Concept Map

Alkanes

Structural properties:

- The general molecular formula is C_nH_{2n+2}.
- Alkanes are saturated hydrocarbons because they contain only single covalent bonds.
- Each alkane member differs from the next by a $-CH_2-$ unit.

Physical properties:

- Alkanes are insoluble in water.
- As the molecular mass of alkanes increases,
  - densities of alkanes increase,
  - melting points of alkanes increase,
  - boiling points of alkanes increase,
  - alkanes become more viscous,
  - alkanes become less flammable.

Chemical properties:

- Combustion
  - Complete combustion gives CO_2 and H_2O.
  - Incomplete combustion also gives C and CO.
- Substitution
  - Alkanes react with chlorine in the presence of UV light to produce a mixture of products.

Exercise 22

Foundation

1. Which of the following represent saturated hydrocarbons?
   1) $CH(CH_3)_3$ 2) CH_2CHCH_3
   3) $CH_3CH(CH_3)CH_3$ 4) $(CH_3)_2CCHCH_3$

   A 1 and 3 only. **B** 2 and 4 only.
   C 1, 2 and 3 only. **D** 1, 2, 3 and 4.

2. When the number of carbon atoms per molecule in alkanes increases, ______________ .
   A the viscosity increases
   B the flammability increases
   C the density decreases
   D the boiling point decreases

3. Which of the following statements about an organic compound with the molecular formula $C_{16}H_{34}$ is correct?
   A It is soluble in water.
   B It is a gas at room temperature.
   C It has the same chemical properties as hexane (C_6H_{14}).
   D It burns in excess oxygen to form carbon monoxide and water.

4. Which of the following statements about alkanes is correct?
   A They undergo addition reactions.
   B They have the same molecular formula.
   C They occur naturally in crude oil.
   D They have the same physical properties.

5. 1000 cm^3 of propane is burnt completely in oxygen.
   a) What is the full structural formula of propane?
   b) Draw the electronic structure of propane.
   c) Write the balanced chemical equation for the complete combustion of propane.

6. The hydrocarbon *J* is an alkane. The relative molecular mass of *J* is 72.
   a) What is meant by 'alkane'?
   b) What is the molecular formula of *J*?
   c) *J* reacts with chlorine in the presence of sunlight.
      i) What type of reaction occurs when *J* reacts with chlorine?
      ii) Write an equation for the reaction between *J* and chlorine.
   d) *J* can also react with bromine vapour in the presence of sunlight. Describe what you would observe when *J* reacts with bromine vapour.

Challenge

1. 50.0 cm^3 of methane is burnt in excess oxygen. What is the volume of carbon dioxide obtained? (All volumes are measured at the same conditions.)
   A 50 cm^3 **B** 75 cm^3
   C 100 cm^3 **D** 150 cm^3

2. When 1.0 mol of an alkane is burnt completely in excess oxygen, 4.0 mol of water molecules are produced. The alkane is ______________ .
   A methane **B** ethane
   C propane **D** butane

3. The table below shows some information about four organic compounds *P*, *Q*, *R* and *S*.

Organic compound	Molecular formula	Melting point (°C)	Boiling point (°C)
P	C_3H_8	–188	–42
Q	C_4H_{10}	–138	–1
R	C_5H_{12}	–130	36
S	C_6H_{12}	6	80

   a) i) Which homologous series does C_3H_8 belong to?
      ii) Why are *P*, *Q*, *R* and *S* classified as hydrocarbons?
   b) Which of these organic compounds belong to the alkane series?
   c) Based on the information given above, state **one** characteristic of the alkane series.
   d) Draw the structures of **two** isomers of *Q*.

Chemistry Today

On December 5, 2002, a small ship collided with a bigger cargo ship in the middle of the Singapore Straits, releasing several tonnes of crude oil. Six anti-pollution crafts were immediately sent out to clean up the oil spill.

Oil spills can be harmful to plants and animals. When birds come in contact with the oil spills, their feathers stick together, leaving them unable to fly. In a cold climate, the birds can even freeze to death as the oil prevents the feathers from keeping the birds warm. If the birds try to clean themselves, they may swallow the oil. The oil will poison them, damaging their lungs, kidneys and livers.

Oil that gets swept up to the mainland can get stuck on plants as well. When oil covers the roots of a plant, the plant will not be able to take in water. If the leaves of the plants get covered with oil, the plant will not be able to make food through photosynthesis as sunlight is blocked.

There are certain strains of bacteria, known as oil-eating bacteria, that can help clean up oil spills. However, this process is slow. Another option is to spray chemicals over the oil spills to dissolve the oil.

CRITICAL THINKING

Apart from endangering plants and animals, what are the other dangers of an oil spill?

Chapter 23

Alkenes

Why is it important to store bananas away from the rest of your fruits and vegetables at home? Ripening bananas produce a gas called ethene. Ethene makes fruits ripen faster and can lead to food spoilage.

Ethene is an organic compound and a hydrocarbon just like ethane, which you have studied about in the previous chapter. However, ethene belongs to another homologous series known as **alkenes**. In this chapter, you will learn about ethene and other alkenes.

Chapter Outline

23.1 | Structure of Alkenes

What are alkenes?

Alkenes are hydrocarbons that contain one or more **carbon-carbon double bonds**. *Any molecule that contains carbon-carbon double bonds* is described as **unsaturated**. Thus, alkenes are called **unsaturated hydrocarbons.**

Chem-Aid

Alkanes are saturated hydrocarbons. They do not contain carbon-carbon double bonds.

What is the functional group of alkenes?

The functional group of alkenes is the carbon-carbon double bond. A carbon-carbon double bond is a *double covalent bond between two carbon atoms* (Fig. 23.1).

Link

What is a double covalent bond? Recall what you have learnt in chapter 7.

carbon-carbon double bond

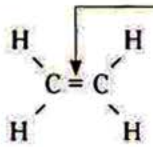


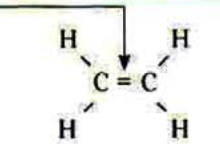
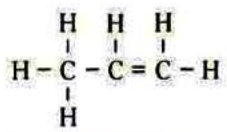
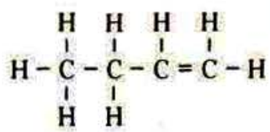
Fig. 23.1 Alkene, an unsaturated hydrocarbon

Table 23.1 shows the various formulae and boiling points of the first three members of the alkene homologous series.

The first member of the alkene series is ethene.

Every alkene contains at least one carbon-carbon double bond.

Ethene, propene and butene are all gases at room temperature.

Name	Molecular formula	Structural formula	Full structural formula	Boiling point (°C)
Ethene	C_2H_4	CH_2CH_2		-104
Propene	C_3H_6	CH_3CHCH_2		-48
Butene	C_4H_8	$CH_3CH_2CHCH_2$		-6

Butene differs from propene by having an additional $-CH_2-$ unit.

The boiling point of an alkene increases as the number of carbon atoms in the molecule increases.

Table 23.1 The first three alkenes

What is the general formula for the alkene homologous series?

Alkenes have the general formula $\mathbf{C_nH_{2n}}$. If an alkene has five carbon atoms per molecule, n = 5. Its molecular formula is $C_5H_{(2\times5)}$, i.e. C_5H_{10}.

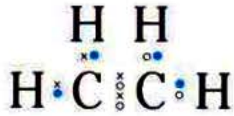


Fig. 23.2(a) Electronic structure of ethene

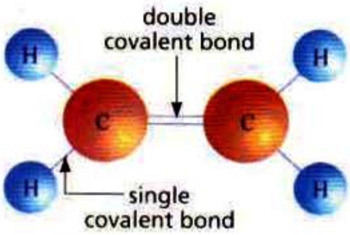


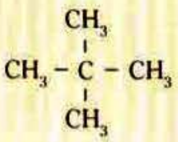
Fig. 23.2(b) Model of an ethene molecule

TidBit

Pentane (boiling point 36 °C) exists as a liquid at room temperature and pressure while its isomer, neopentane (boiling point 9 °C) is a gas. The structural formulae of pentane and neopentane is shown below.

$CH_3 - CH_2 - CH_2 - CH_2 - CH_3$

Pentane



Neopentane

What is the electronic structure of ethene?

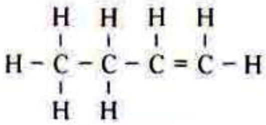
The ethene molecule contains a carbon-carbon double bond. This covalent bond is formed by the sharing of two pairs of electrons between the carbon atoms (Fig. 23.2).

Isomerism in Alkenes

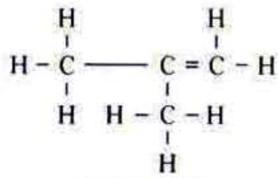
Consider the two hydrocarbons A and B below.

What similarities and differences are there between hydrocarbons A and B?

Both A and B have the same molecular formula, C_4H_8, but they have different structural formulae. Thus, A and B are isomers. A is called butene while B is called methylpropene.



Hydrocarbon A

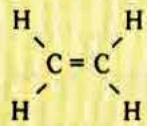


Hydrocarbon B

Butene is a straight-chain unsaturated hydrocarbon while methylpropene is a branched-chain unsaturated hydrocarbon. Butene and methylpropene have different melting and boiling points.

Key ideas

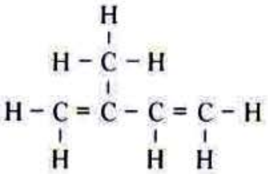
1. Alkenes are unsaturated hydrocarbons. They contain one or more carbon-carbon double bonds.
2. The functional group of alkenes is the carbon-carbon double bond.
3. The general formula of alkenes is C_nH_{2n}.
4. The first member of the alkene homologous series is ethene (C_2H_4).
5. Alkenes can exist as branched or unbranched hydrocarbons.
6. Isomers have the same molecular formula but different structural formulae.



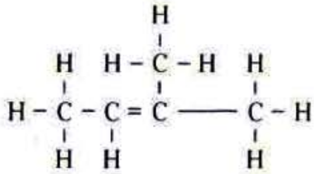
Test Yourself 23.1

Worked Example

The structures of two organic compounds, *A* and *B*, are shown.



Compound *A*



Compound *B*

Which of the following statements are correct?
1) They are both hydrocarbons.
2) They are both alkenes.
3) They are isomers of each other.

A 1 and 2 only. B 1 and 3 only.
C 2 and 3 only. D 1, 2 and 3.

Thought Process
- Hydrocarbons are compounds containing carbon and hydrogen only.
- Alkenes are hydrocarbons containing at least one C = C double bond.
- Isomers are organic compounds containing the same molecular formula, but different structural formulae.

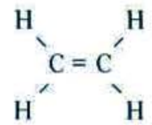
The molecular formula of compound *A* is C_5H_8. The molecular formula of compound *B* is C_5H_{10}. They are alkenes but they are not isomers of each other.

Answer
A

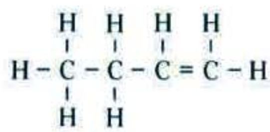
Questions
1. What are the differences between an ethene molecule and an ethane molecule?
2. Draw the electronic structure of a propene molecule.
3. Consider the molecular models of the two organic compounds shown below.
   a) Write the molecular formulae of compounds I and II.
   b) Explain why compounds I and II are **not** isomers.
   c) State **one** similarity between these two compounds in terms of chemical bonding.

TidBit

Many plants produce alkenes. These volatile hydrocarbons may give the plants a distinct aroma. For example, oranges and lemons get their characteristic smell from an alkene called limonene.



Compound I

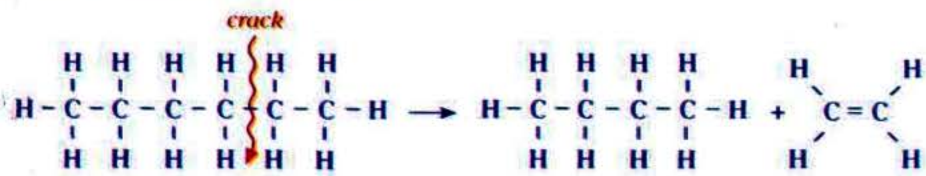


Compound II

23.2 | Manufacturing Alkenes by Cracking

Alkenes are obtained by **cracking** petroleum (crude oil). Cracking is *the breaking down of long-chain hydrocarbons into smaller molecules*. For example,

hexane ⟶ butane + ethene



Chem-Aid

A catalyst is a substance that speeds up a reaction but remains unchanged at the end of the reaction.

A **catalyst** may be used to speed up the process of cracking. This process is known as **catalytic cracking**.

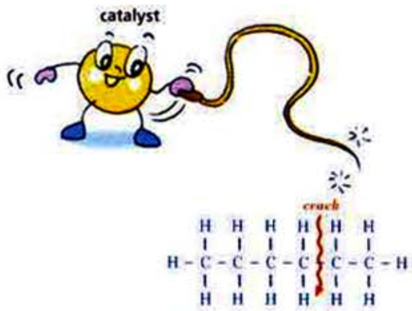
Link

Petroleum is a mixture of alkanes. Recall what you have learnt in chapter 21.

In industries, cracking is done by passing the petroleum fraction containing long chains of carbon atoms over a catalyst at a high temperature (about 600 °C). The catalyst is either aluminium oxide or silicon(IV) oxide. The following reaction takes place.

$$\text{long-chain alkanes} \xrightarrow{\text{catalytic cracking}} \text{mixture of short-chain alkenes} + \text{mixture of short-chain alkanes or hydrogen gas}$$

Cracking converts saturated hydrocarbons (alkanes) into unsaturated hydrocarbons (alkenes). For example, on cracking, naphtha (a mixture of long-chain alkanes) becomes C_2 to C_4 alkenes. C_2 to C_4 alkenes refer to ethene, propene and butene. These alkenes are used to produce petrochemicals.



Cracking splits long-chain alkanes into short-chain alkanes and alkenes.

The products obtained on cracking will depend on the conditions used. However, an alkene is always produced. Hydrogen gas may also be produced. This hydrogen gas can be used to manufacture ammonia in the Haber process (see chapter 19).

Catalytic cracking can be demonstrated in the school laboratory using the apparatus as shown in Fig. 23.3. Petroleum vapour is passed over a heated porous pot. The material of the porous pot contains the catalysts aluminium oxide and silicon(IV) oxide. The product is a mixture of alkanes and alkenes especially ethene.

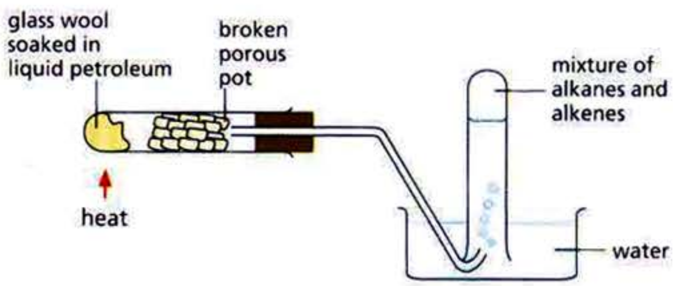


Fig. 23.3 Cracking liquid petroleum in the school laboratory

Why is cracking important?

Cracking is a very important process for three reasons:

1. **Cracking is used to produce petrol.**
   Our need for petrol is greater than our need for diesel oil or lubricating oil. Diesel and lubricating oils are made up of long-chain alkanes. Through cracking, these long-chain molecules can be converted into petrol (short-chain alkanes).

2. **Cracking is used to produce short-chain alkenes.**

$$\text{long-chain alkane} \longrightarrow \text{shorter-chain alkane} + \text{alkene}$$

$$C_{18}H_{38} \longrightarrow C_6H_{14} + 6C_2H_4$$

Short-chain alkenes such as ethene and propene are used as starting materials for making ethanol and plastics.

3. Cracking is used to produce hydrogen.

$$C_{18}\text{ alkane molecule} \longrightarrow C_8\text{ alkene molecule} + C_{10}\text{ alkene molecule} + \text{hydrogen}$$

$$C_{18}H_{38} \longrightarrow C_8H_{16} + C_{10}H_{20} + H_2$$

Hydrogen is a by-product in the cracking of alkanes.

TidBit

This is a plant for catalytic cracking. Such plants are often known as 'cat-cracker' plants.

23.3 | Chemical Properties of Alkenes

Alkenes take part in **combustion** and **addition** reactions.

Combustion

Alkenes burn in a plentiful supply of air (oxygen) to form carbon dioxide and water vapour.

$$\text{ethene} + \text{oxygen} \xrightarrow{\text{heat}} \text{carbon dioxide} + \text{water vapour}$$

$$C_2H_4(g) + 3O_2(g) \xrightarrow{\text{heat}} 2CO_2(g) + 2H_2O(g)$$

Since alkenes have a relatively higher percentage of carbon than the corresponding alkanes, carbon particles may be produced on combustion. Hence, alkenes burn with a smokier flame than alkanes with a similar number of carbon atoms.

Addition Reactions

Why do alkenes undergo addition reactions?
The carbon-carbon double bonds in alkenes are reactive and thus will readily undergo **addition reactions**.

The carbon-carbon double bond makes butene reactive.

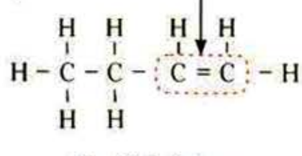
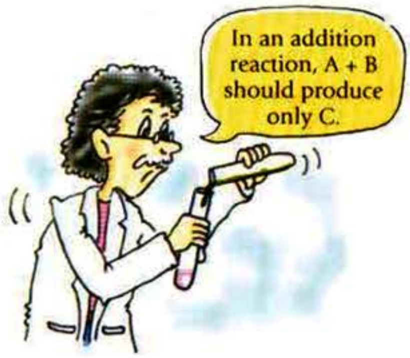


Fig. 23.4 Butene

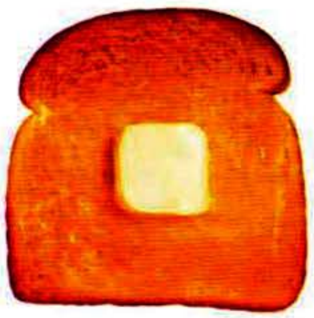
What is an addition reaction?
An addition reaction is *a reaction in which an unsaturated organic compound combines with another substance to form a single new compound.* Addition reactions of alkenes can be represented by the general equation

In an addition reaction, carbon-carbon double bonds become single bonds. Hence, an unsaturated hydrocarbon becomes a saturated organic compound.



We shall examine three addition reactions using ethene as an example:

1. **Hydrogenation** — addition of hydrogen to alkenes
2. **Bromination** — addition of bromine to alkenes
3. **Hydration** — addition of steam to alkenes



Thanks to hydrogenation, we can have margarine on toast.

Hydrogenation – Addition of Hydrogen to Alkenes

At 200 °C and in the presence of a catalyst such as nickel, alkenes can react with hydrogen to form alkanes. This reaction is known as **hydrogenation**.

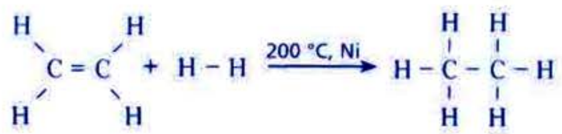
alkene + hydrogen ⟶ alkane

$$C_nH_{2n} + H_2 \longrightarrow C_nH_{2n+2}$$

For example,

ethene + hydrogen $\xrightarrow{200\,°\text{C, Ni}}$ ethane

$$C_2H_4(g) + H_2(g) \xrightarrow{200\,°\text{C, Ni}} C_2H_6(g)$$



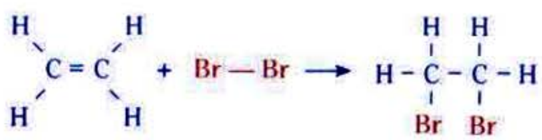
How is hydrogenation useful to us?
Hydrogenation is used to make **margarine** from vegetable oil. You will learn more about margarine in section 23.5.

Bromination – Addition of Bromine to Alkenes

A solution of bromine is reddish-brown (when concentrated) or yellow (when dilute). If an alkene is passed into or added to a solution of bromine, the yellow or reddish-brown colour disappears (decolourises) *immediately* and a colourless oil is formed. This reaction is called **bromination.** For example,

ethene + bromine ⟶ 1,2-dibromoethane

$$C_2H_4(g) + Br_2(aq) \longrightarrow C_2H_4Br_2(l)$$



How is bromination useful to us?

This reaction serves as a *chemical test for the presence of unsaturated hydrocarbons*. It is used to distinguish between an alkane and an alkene. Alkanes do not decolourise bromine solution under normal conditions, but alkenes decolourise bromine solution rapidly (Fig. 23.5).

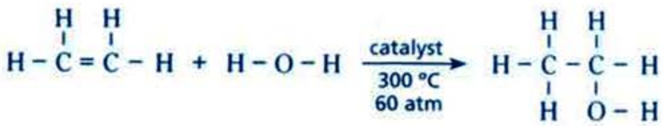
Fig. 23.5 Testing for the presence of unsaturated hydrocarbons

Hydration — Addition of Steam to Alkenes

Ethene can react with steam to produce ethanol, which is an alcohol. The reaction requires a catalyst (phosphoric(V) acid). Other conditions required are a temperature of 300 °C and a pressure of 60 atm.

$$\text{ethene} + \text{steam} \xrightarrow[\text{300 °C, 60 atm}]{\text{catalyst}} \text{ethanol}$$

$$C_2H_4(g) + H_2O(g) \xrightarrow[\text{300 °C, 60 atm}]{\text{catalyst}} C_2H_5OH(l)$$



Link

Ethene molecules can also react with each other to form a long molecule called poly(ethene). Find out more in chapter 25.

23.4 | Comparing Alkanes and Alkenes

Below are some similarities and differences between alkanes (saturated hydrocarbons) and alkenes (unsaturated hydrocarbons).

Similarities

1. Both alkanes and alkenes are compounds that contain only carbon and hydrogen.
2. Both alkanes and alkenes are flammable. On complete combustion, they form carbon dioxide and water.

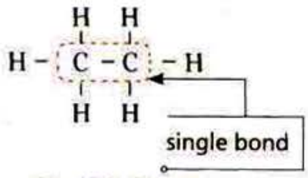


Fig. 23.6 Ethane (saturated hydrocarbon)

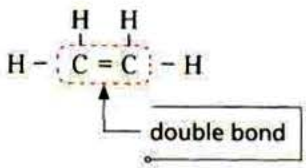


Fig. 23.7 Ethene (unsaturated hydrocarbon)

Differences

1. **Molecular structure**
   Alkanes contain only single bonds between carbon atoms, whereas alkenes contain double bonds between carbon atoms.

2. **Reactivity**
   Alkanes are generally unreactive but alkenes are very reactive.

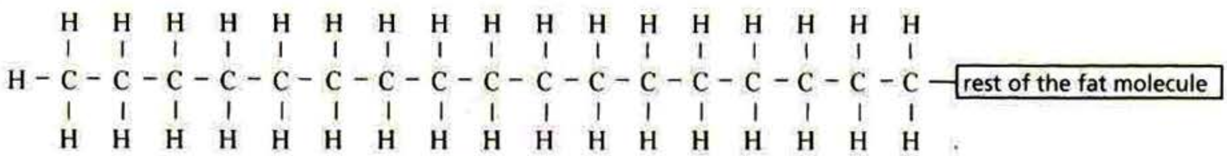
3. **Type of reaction**
   Alkanes undergo substitution reactions but alkenes undergo addition reactions. Furthermore, alkanes do not undergo polymerisation, whereas alkenes undergo addition polymerisation.

4. **Reaction with aqueous bromine**
   Alkanes do not react with aqueous bromine under normal conditions but alkenes rapidly decolourise aqueous bromine.

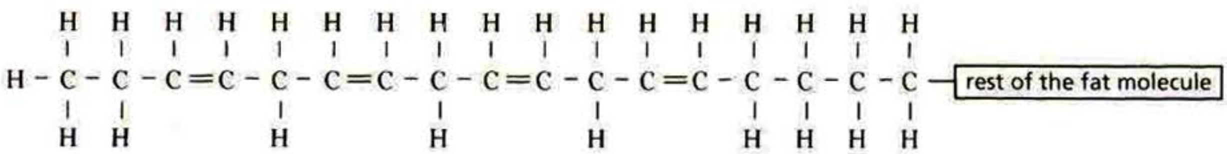
5. **Combustion**
   Alkenes produce a smokier flame than alkanes with a similar number of carbon atoms.

23.5 | Fats and Oils

You may be familiar with fats and oils such as butter, lard and corn oil. There are actually two types of fats: unsaturated fats and saturated fats. The structures of both types of fats are shown below.



Molecule A



Molecule B

Molecule A represents a saturated fat because it consists only of carbon atoms with single bonds between them. Molecule B is an unsaturated fat because double bonds exist between the carbon atoms.

Some fats and oils are called **polyunsaturated** fats and oils. This is because their hydrocarbon chains contain more than one carbon-carbon double bond.

Chem-Aid

The carbon-carbon double bonds in unsaturated fat molecules give them a 'bent' structure.

What are the differences between fats and oils?

Fats are solids at room temperature and pressure. They contain mainly saturated fat molecules. Oils are liquids at room temperature and pressure. They contain a larger proportion of unsaturated fat molecules.

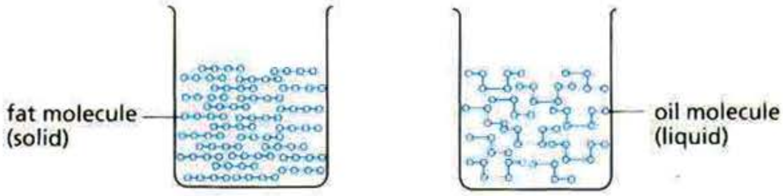


Fig. 23.8 The arrangements of fat molecules and oil molecules. Fat molecules are straight and thus can be packed closely, forming a solid.

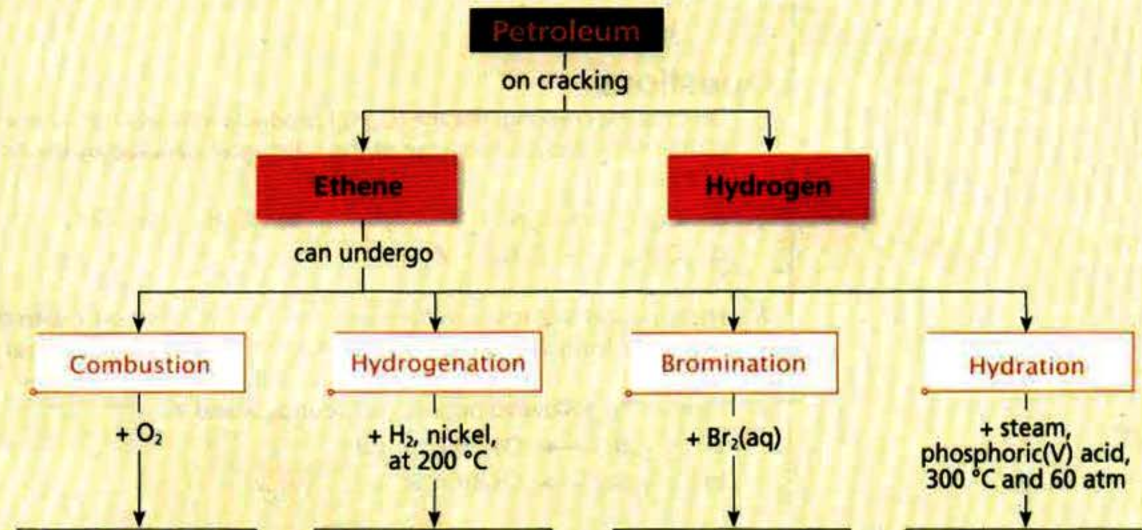
How is margarine produced?

To produce margarine, we add hydrogen to vegetable oil. A temperature of 200 °C and a nickel catalyst are required. The greater the amount of hydrogen used, the more solid the margarine becomes.

Margarine

Key ideas

1. The cracking of alkanes produces alkenes and smaller alkanes. Hydrogen may also be produced.
2. The reactions of ethene are as follows:

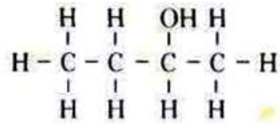


3. Polyunsaturated fats and oils contain many carbon-carbon double bonds.
4. Hydrogenation is used to convert vegetable oils to margarine.

Test Yourself 23.2

Worked Example

Compound *Y* reacts with steam to form compound *Z* with the structural formula shown on the right.

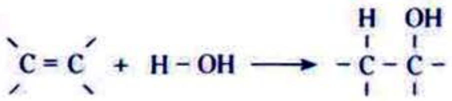


Identify *Y*.

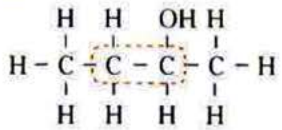
Thought Process

Compound *Z* is an alcohol. An alkene reacts with steam to form an alcohol. Thus, compound *Y* is an alkene.

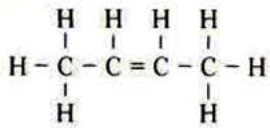
The reaction between an alkene and steam can be represented by the following equation:



H and OH are added across these two carbon atoms. The C = C bond must be located here.



Answer

Y is 

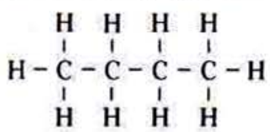
Questions

1. On catalytic cracking, decane ($C_{10}H_{22}$) produces a number of substances. Write the molecular formulae of the substances denoted by the letters *X*, *Y* and *Z*.
   a) $C_{10}H_{22} \longrightarrow C_2H_4 + X$
   b) $C_{10}H_{22} \longrightarrow C_7H_{16} + Y$
   c) $C_{10}H_{22} \longrightarrow C_{10}H_{20} + Z$

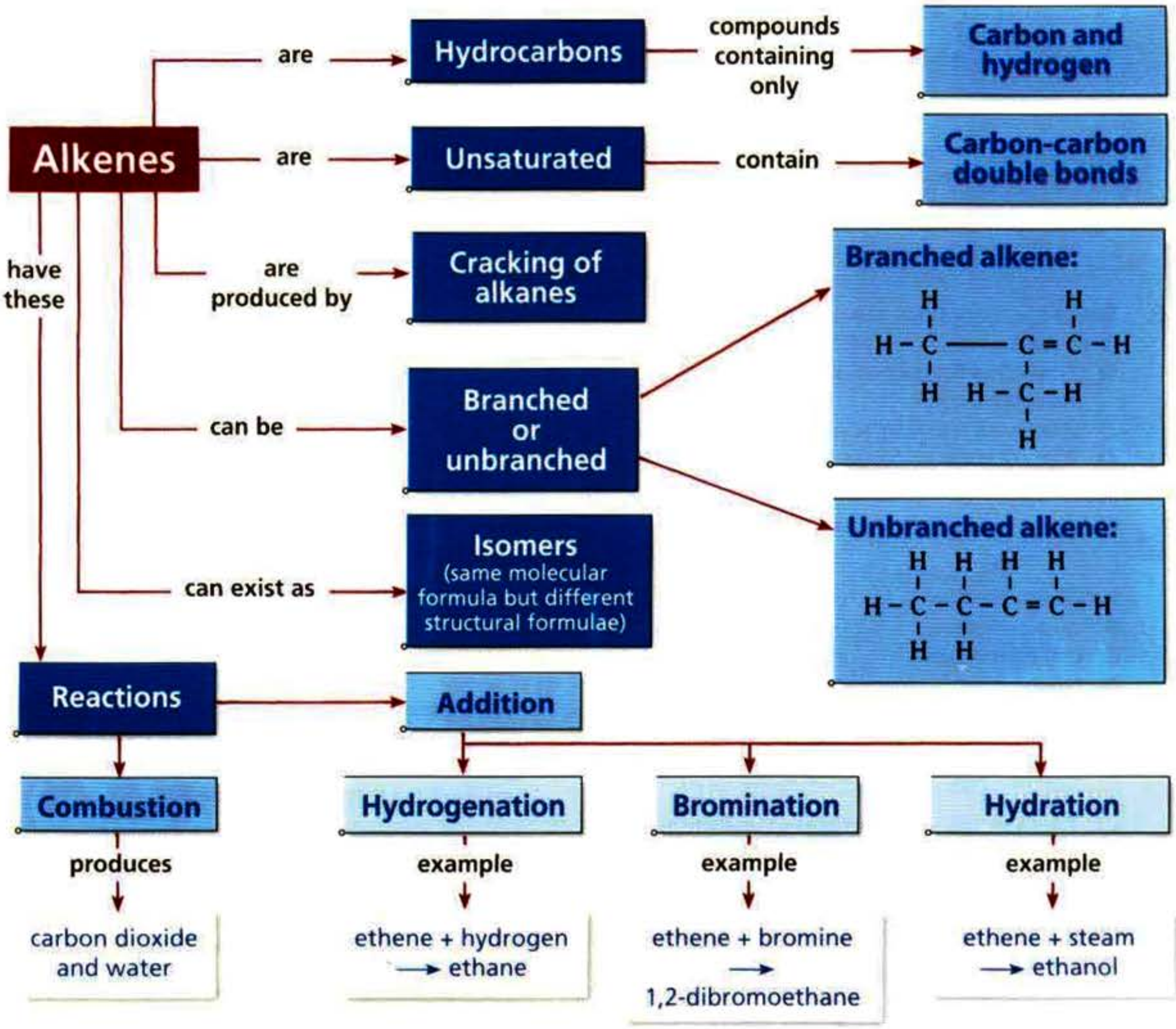
2. Hydrocarbon *X* reacts with hydrogen to form hydrocarbon *Y* with the molecular formula C_3H_8. Draw the full structural formulae of *X* and *Y*.

3. Identify the following organic compounds, *X* and *Y*.
   a) $X + Br_2 \longrightarrow CH_3CH_2Br + HBr$
   b) $Y + Br_2 \longrightarrow CH_2BrCH_2Br$

4. Draw the structure of two alkenes that can be hydrogenated to make the following organic compound.



Concept Map



Exercise 23

Foundation

1. The structural formulae of four hydrocarbons are shown below.

   1) $CH_3CH=CHCH_3$

   2) $\begin{array}{c} CH_3-C-CH_3 \\ \| \\ CH_2 \end{array}$

   3) $CH_2=CHCH_2CH_3$

   4) $\begin{array}{c} CH_2-CH_2 \\ | \quad\quad | \\ CH_2-CH_2 \end{array}$

   Which structural formulae represent alkenes?

   A 1 and 2 only.
   B 2 and 3 only.
   C 1, 2 and 3 only.
   D 1, 2, 3 and 4.

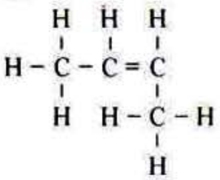
2. Both ethene and ethane have the same ______.

   A density
   B general formula
   C chemical properties
   D number of carbon atoms

3. Hexane has the formula C_6H_{14}. Which of the following cannot be formed by cracking hexane?

   A H_2
   B C_2H_4
   C C_4H_{10}
   D C_7H_{16}

4. The structural formula of a hydrocarbon is shown below.



What product is formed when the hydrocarbon reacts with hydrogen?

A An alkane.
B A polymer.
C An alcohol.
D A carboxylic acid.

5. The properties of two gases, *P* and *Q*, are given here.
   1) *P* dissolves in aqueous sodium hydroxide but *Q* does not.
   2) *Q* does not decolourise aqueous bromine.
   3) When *Q* burns in excess oxygen, the products are *P* and water only.

   What are *P* and *Q* likely to be?

	P	*Q*
A	carbon dioxide	ethane
B	carbon dioxide	ethene
C	carbon monoxide	ethane
D	ethane	ethene

6. a) What is the general formula of alkenes?
   b) Describe how alkenes are manufactured.
   c) The table shows the properties of some alkenes.

Alkene	Molecular formula	Boiling point (°C)
ethene	C_2H_4	−104.0
	C_3H_6	−47.7
butene		−6.2
pentene	C_5H_{10}	30.0
hexene	C_6H_{12}	64.0

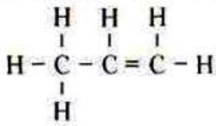
   i) Complete the above table.
   ii) Describe the trend of the boiling points of alkenes.
   iii) Draw a 'dot and cross' diagram to show the bonding in ethene. (You only need to show the valence electrons.)
   iv) What is the value of *x* for an alkene with the molecular formula, C_xH_{16}?

   d) i) Identify the compound *X* in the following reaction:

   ethene + *X* ⟶ ethanol

   ii) Using full structural formulae, write the equation for the reaction.
   iii) What name is given to this type of reaction?

7. a) Name the hydrocarbon shown below.



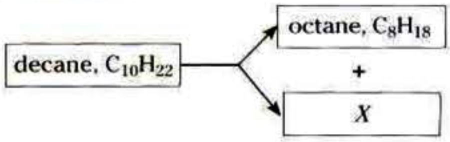
   b) Is this molecule unsaturated? Explain your answer.
   c) Write the chemical equation to represent the complete combustion of this hydrocarbon in an excess of oxygen.
   d) i) Name the homologous series to which this hydrocarbon belongs.
   ii) Draw the full structural formula of the hydrocarbon with one more carbon atom in the same homologous series.

8. When hexane (C_6H_{14}) was heated with a catalyst, two hydrocarbon gases, *A* and *B*, were formed. Both gases were then tested with bromine water. The results of the tests are shown below.

Hydrocarbon gas *A*	Hydrocarbon gas *B*
Bromine water remains the same colour.	Bromine water turns colourless.

   a) What type of chemical change will occur when hexane was heated with this catalyst?
   b) What can you deduce from the results of the tests with bromine water?

9. Below is a flow diagram showing the cracking of an alkane.



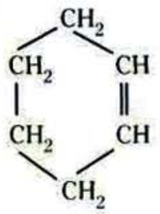
   a) Name substance *X*.
   b) Describe how alkanes are cracked.

Challenge

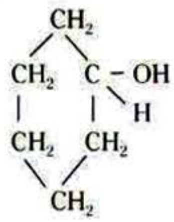
1. Which of the following compounds can undergo an addition reaction with chlorine?

   A $CH_3(CH_2)_4CH_3$ **B** CH_3CH_2OH

   C

   D

2. One mole of a hydrocarbon *X* reacts with one mole of hydrogen to form a saturated hydrocarbon. What could be the formula of *X*?

   A C_3H_8 **B** C_6H_{12}

   C C_4H_{10} **D** C_7H_{16}

3. When an alkene reacts with bromine, the product formed is

$$\begin{array}{ccccccccc} & & H & & Br & & H & & H & & \\ & & | & & | & & | & & | & & \\ Br & - & C & - & C & - & C & - & C & - & Br \\ & & | & & | & & | & & | & & \\ & & H & & H & & Br & & H & & \end{array}$$

   This implies that the number of carbon-carbon double bonds in this alkene is ________.

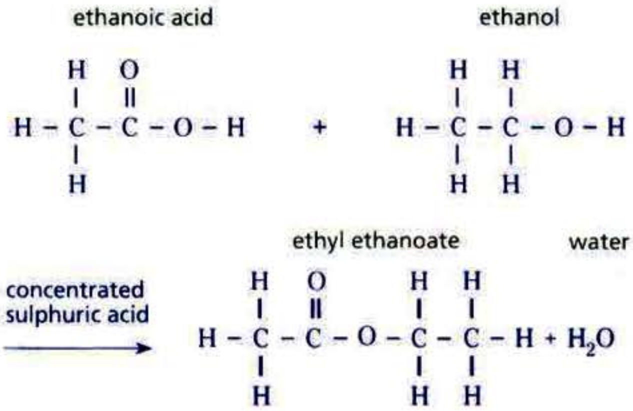
   A one **B** two
   C three **D** four

How do we prepare esters?

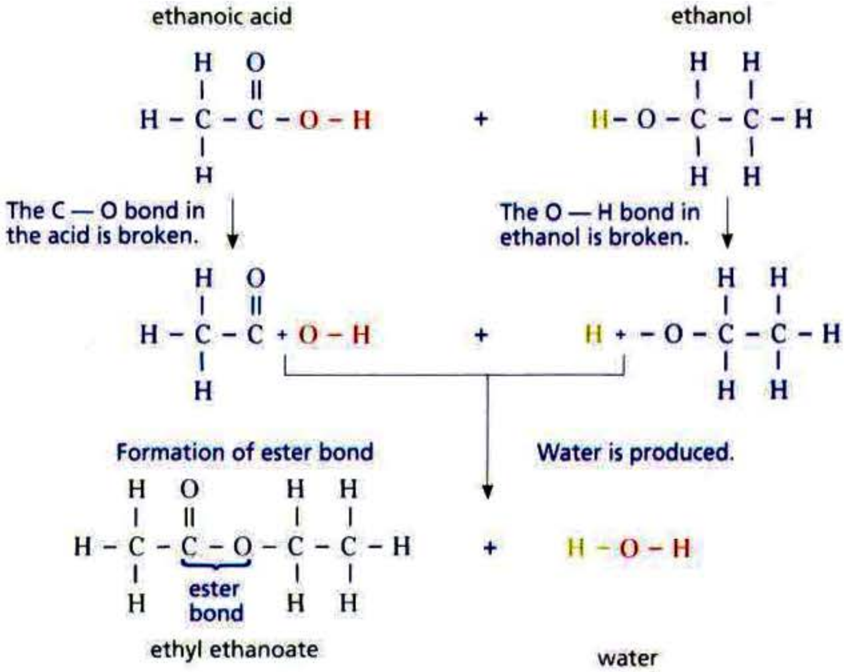
If ethanoic acid is warmed with ethanol in the presence of a few drops of concentrated sulphuric acid (Fig. 24.7), an ester called ethyl ethanoate is formed. Concentrated sulphuric acid acts as a catalyst to speed up the reaction.

Link

Esterification is used to make the macromolecule Terylene. Read more in chapter 25.



This is what happens when ethanoic acid reacts with ethanol during the formation of the ester:



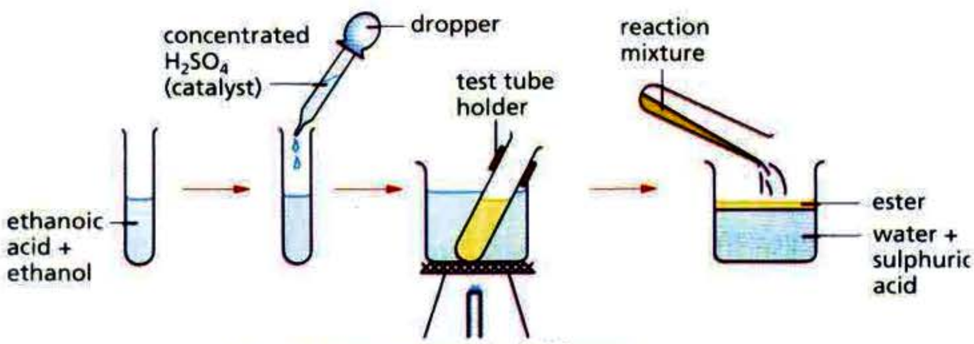


Fig. 24.7 Making an ester in the laboratory

Chapter 24

Alcohols and Carboxylic Acids

Chapter Outline

It is dangerous to drink alcohol and then drive. But is it possible to use alcohol as a fuel in your car? Believe it or not, cars running the famous Indianapolis 500 race are only allowed to use alcohol as fuel. Is this the same type of alcohol that is commonly consumed? What are the advantages of using alcohol as a fuel? What reaction takes place when alcohol combusts? In this chapter, you will find out the answers to these questions and more.

24.1 | Structure of Alcohols

What are alcohols?

In chapters 22 and 23, you studied two homologous series: the alkanes and the alkenes. The alcohols are another homologous series. Alcohols are organic compounds which have the **hydroxyl** (–OH) functional group (Fig. 24.1).

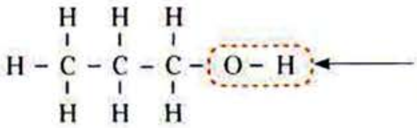

All alcohols have this functional group

Fig. 24.1 The functional group, –OH, in alcohols

All alcohols contain the elements carbon, hydrogen and oxygen. They have the general formula $\mathbf{C_nH_{2n+1}OH}$.

Naming Alcohols

The name of an alcohol ends with '-ol'. Hence, an alcohol that contains two carbon atoms is called ethanol. The first four alcohols in the homologous series are given in Table 24.1.

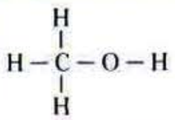
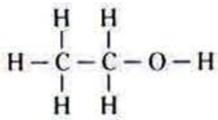
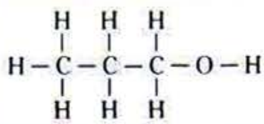
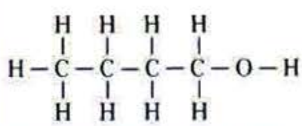
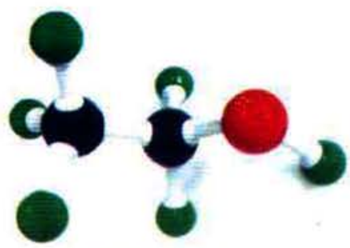
Name	Molecular formula	Structural formula	Full structural formula
Methanol	CH_3OH or CH_4O	CH_3OH	
Ethanol	C_2H_5OH or C_2H_6O	CH_3CH_2OH	
Propanol	C_3H_7OH or C_3H_8O	$CH_3CH_2CH_2OH$	
Butanol	C_4H_9OH or $C_4H_{10}O$	$CH_3CH_2CH_2CH_2OH$	

Table 24.1 The first four alcohols

TidBit

Perfumes often contain rose oil. An organic compound called geraniol gives rose oil its distinct smell. Geraniol contains two functional groups, namely, the carbon-carbon double bond and the hydroxyl group.


A 3-D model of ethanol

24.2 | Properties of Alcohols

Try it Out

The alkanes CH_4 to C_4H_{10} and the alkenes C_2H_4 to C_4H_8 are gases at room temperature and pressure. Why are alkanes and alkenes gases, when alcohols are volatile liquids at room temperature and pressure? You may use the Internet to find out.

Table 24.2 shows the physical state, solubility and boiling points of some alcohols.

Alcohol	Physical state	Solubility in water	Boiling point (°C)
Methanol	liquid	very soluble	65
Ethanol	liquid	very soluble	78
Propanol	liquid	soluble	97
Butanol	liquid	slightly soluble	118

Table 24.2 The solubilities and boiling points of some alcohols

- Alcohols are soluble in water, but their solubility decreases as the molecular size increases. For example, methanol is very soluble in water but butanol is only slightly soluble in water.
- Unlike the alkanes and alkenes, the first four alcohols are liquids at room temperature and pressure.

Chemical Properties of Alcohols

- Although alcohols contain the –OH group, they are not alkalis. In fact, they are all neutral.
- Alcohols are more reactive than alkanes because the C–O and O–H bonds in alcohols are more reactive than the C–C and C–H bonds in alkanes.

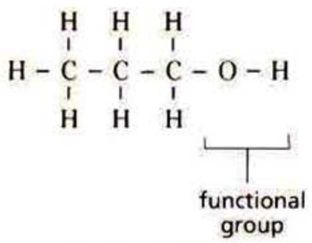


Fig. 24.2 The functional group (–OH) of alcohols is responsible for the typical reactions that alcohols undergo.

- All alcohols have similar chemical properties. Alcohols can take part in these reactions: combustion and oxidation.

Combustion

Like most other organic compounds, an alcohol such as ethanol, burns in air to produce carbon dioxide and water vapour.

$$\text{ethanol} + \text{oxygen} \xrightarrow{\text{heat}} \text{carbon dioxide} + \text{water vapour}$$

$$C_2H_5OH(l) + 3O_2(g) \xrightarrow{\text{heat}} 2CO_2(g) + 3H_2O(g)$$

The combustion of alcohols in excess oxygen produces a clean flame, as only carbon dioxide and water vapour are produced.

How is the combustion of alcohols useful to us?
Alcohols can be used as a fuel. In fact, some race cars run on methanol. Methanol is less volatile than petrol and is less likely to explode in an accident. Methanol is also a clean fuel. It does not produce soot on combustion. Singapore relies heavily on petrol as a fuel. Do you think it is good to use alcohol as a fuel instead?

Alcohol is also burnt on some foods such as fruit cake to give it a distinct flavour.

Oxidation

Alcohols are easily oxidised. For example, in the laboratory, we can oxidise ethanol by heating it with a mixture of potassium dichromate(VI) solution ($K_2Cr_2O_7$) and dilute sulphuric acid (Fig. 24.3). Ethanol then becomes **ethanoic acid**.

Chem-Aid

A mixture of potassium dichromate(VI) and sulphuric acid is known as 'acidified potassium dichromate(VI)'.

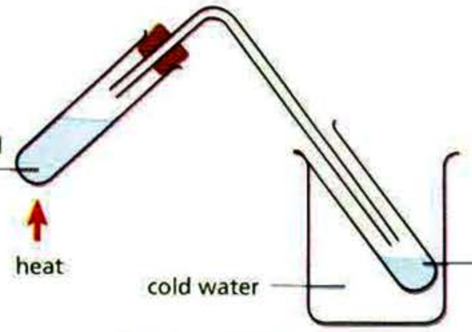


Fig. 24.3 The oxidation of ethanol

Link

Recall what you have learnt about redox reactions in chapter 13.

ethanol + oxygen from oxidising agent → ethanoic acid + water

$$\begin{array}{c} \text{H}\;\;\text{H} \\ |\;\;\;\;| \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{H} \\ |\;\;\;\;| \\ \text{H}\;\;\text{H} \end{array} + 2[\text{O}] \longrightarrow \begin{array}{c} \text{H}\;\;\text{O} \\ |\;\;\;\;\| \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{H} \\ | \\ \text{H} \end{array} + \text{H}_2\text{O}$$

Quick Check

On oxidation, ethanol becomes ethanoic acid and propanol becomes propanoic acid. What is the product when methanol is oxidised?

In this redox reaction,

a) ethanol acts as a reducing agent while acidified potassium dichromate(VI) acts as an oxidising agent,
b) potassium dichromate(VI) is reduced. Its colour changes from orange to green.
c) ethanol is oxidised to ethanoic acid. Ethanoic acid is an example of a carboxylic acid. You will learn about carboxylic acids in sections 24.4 – 24.5.

How is the oxidation of alcohols useful to us?
In Singapore, the police use **breathalysers** to test the amount of alcohol consumed by drivers. A breathalyser contains acidified potassium dichromate(VI). If the breath of a driver contains a high level of alcohol, a colour change is registerec on the device.

A breathalyser is used to test the amount of alcohol consumed.

24.3 | Manufacturing Ethanol

Ethanol can be manufactured from
- hydration of ethene (see chapter 23),
- fermentation of carbohydrates.

Manufacturing Ethanol from Ethene

Ethanol is manufactured by the catalytic addition of steam to ethene. The mixture of ethene and steam is passed through phosphoric(V) acid at 300 °C and 60 atm. Phosphoric(V) acid (H_3PO_4) acts as a catalyst in this reaction.

The equation for the reaction between ethene and steam is

$$\text{ethene} + \text{steam} \xrightarrow[300\ ^\circ\text{C},\ 60\ \text{atm}]{H_3PO_4} \text{ethanol}$$

$$CH_2=CH_2(g) + H_2O(g) \xrightarrow[300\ ^\circ\text{C},\ 60\ \text{atm}]{H_3PO_4} CH_3CH_2OH$$

Producing Ethanol by Fermentation

To most people, alcohol refers to beverages like wine and beer. Actually, such drinks contain ethanol. To produce ethanol, we **ferment** fruits, vegetables or grains.

What is fermentation?

Fermentation is a *chemical process in which microorganisms* such as yeast *act on carbohydrates to produce ethanol and carbon dioxide*. Sugars and starch are examples of carbohydrates.

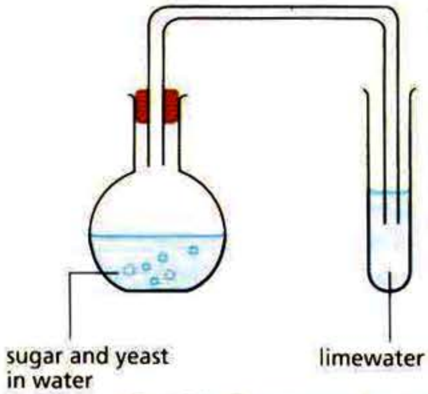
Yeast contains **enzymes** (biological catalysts) which cause starch or sugars to break down to glucose. The glucose is then broken down to ethanol and carbon dioxide.

A brewery where alcohol is produced

Fig. 24.4 shows the apparatus used for fermentation in the laboratory. A glucose solution is mixed with yeast and the mixture is kept at a temperature of about 37 °C. Ethanol and carbon dioxide are produced after a few days.

$$\text{glucose solution} \xrightarrow{\text{yeast}} \text{ethanol} + \text{carbon dioxide}$$

$$C_6H_{12}O_6(aq) \xrightarrow{\text{yeast}} 2C_2H_5OH(aq) + 2CO_2(g)$$



1 In fermentation, carbon dioxide is produced. Hence, **frothing** (formation of foam) can be observed in the flask. A white precipitate can be also observed to form in the limewater.

2 Fermentation can only take place in the absence of oxygen. Hence, the apparatus needs to be air-tight. This is ensured by the following precautions:
- The rubber bung is secured tightly to the flask.
- The limewater in the test tube prevents air from entering the apparatus.

Fig. 24.4 The apparatus for fermentation of glucose solution

The enzymes in yeast work best at around 37 °C. If the temperature is raised beyond 37 °C, the enzymes will die and fermentation stops.

The fermentation of sugars produces only a dilute solution of ethanol (up to about 15%). This is because when the alcohol content exceeds this value, the yeast dies and fermentation stops. Ethanol can be obtained from this liquid mixture by fractional distillation.

In Brazil, ethanol is manufactured by the fermentation of cane sugar. During fermentation, the sucrose ($C_{12}H_{22}O_{11}$) in the cane sugar is first converted to glucose, and then to ethanol and carbon dioxide. This ethanol is used as a fuel in cars!

What happens if alcoholic drinks are left exposed to air?
If an alcoholic drink such as wine or beer is left exposed to the air, it will turn sour after a few days. This is caused by the action of **bacteria** from the air. Using oxygen from the air, the bacteria oxidises ethanol into ethanoic acid.

$$\text{ethanol} + \text{oxygen from air} \xrightarrow{\text{bacteria}} \text{ethanoic acid} + \text{water}$$

$$C_2H_5OH(aq) + O_2(g) \xrightarrow{\text{bacteria}} CH_3COOH(aq) + H_2O(l)$$

It is therefore important that wine is kept tightly corked. This is also the reason why, when we are fermenting sugars to produce alcohol, it is important to ensure that fermentation takes place in the absence of air. Otherwise, the alcohol will be oxidised to form an acid called a carboxylic acid.

Drink	Alcohol content (%)
Beer	3 – 7
Wine	12 – 15
Whisky	40
Brandy	40

Table 24.3 Amounts of alcohol in various drinks

Uses of Ethanol

Ethanol is used in the following ways: in alcoholic drinks, as a solvent, and as a fuel.

In alcoholic drinks

Large quantities of ethanol are used in alcoholic drinks. Table 24.3 shows the amounts of alcohol in various drinks.

Alcohol is a drug. Drinking alcohol can be addictive and can cause serious health problems such as damage to the liver and kidneys, coma or even death.

As a solvent

Ethanol is a good solvent. It dissolves many substances that are not soluble in water. It also evaporates quickly. Hence, it is used in paints, varnishes, deodorants, perfumes, colognes and after-shave lotions. Shellac dissolved in ethanol is used as a liquid polish.

As a fuel

Ethanol can be used as a fuel and is the main constituent of methylated spirit. **Methylated spirit** contains 85% ethanol, 5% methanol and 10% water. Methanol is added to make it unsuitable for drinking. This process is called **denaturing**. Methylated spirit is used in spirit lamps and burners for cooking. For household use, a purple or blue dye is also added to methylated spirit so that it is not confused with other liquids.

In Brazil, ethanol is mixed with petrol and used as fuel to run motor vehicles.

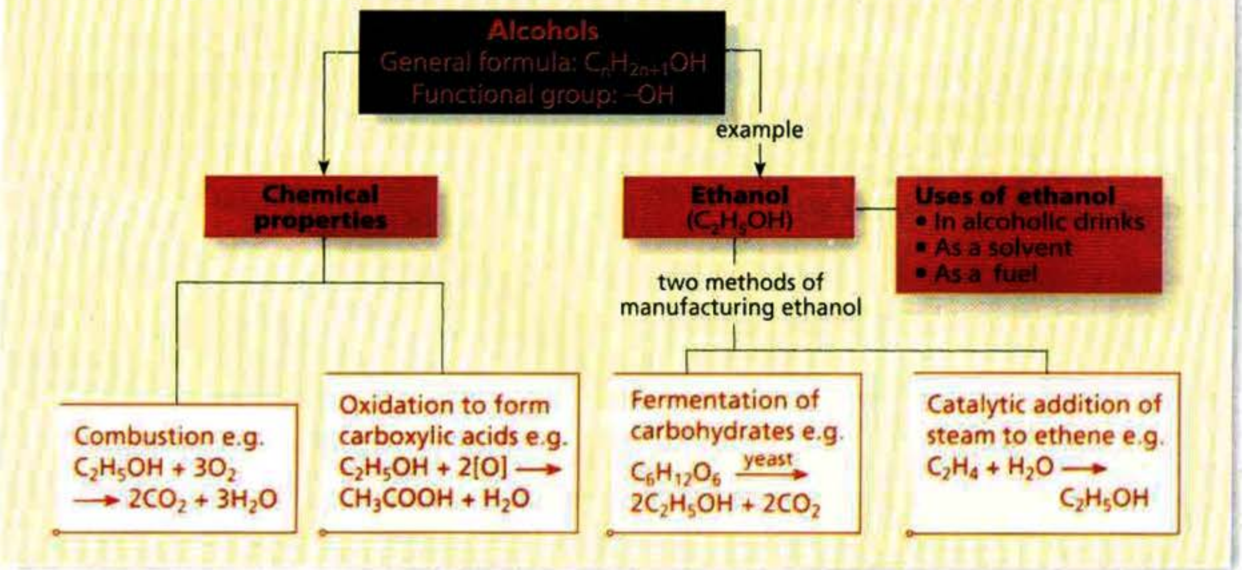
TidBit

Methanol and ethanol have similar chemical properties, but methanol is extremely toxic. Drinking methanol can lead to blindness and even death.

Quick Check

Name an organic substance that is suitable for removing
a) blood stains.
b) ink from a ballpoint pen.

Key ideas



Test Yourself 24.1

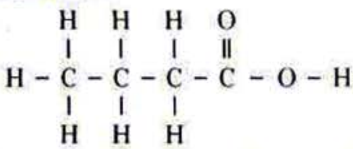
Worked Example

1.0 mol of a straight chain (unbranched) alcohol *X* produces 4.0 mol of carbon dioxide on combustion. Draw the structural formula of the carboxylic acid produced when *X* is oxidised.

Thought Process

1.0 mol of the alcohol produces 4.0 mol of carbon dioxide. Thus, the alcohol contains four carbon atoms. The alcohol is butanol. When we oxidise an alcohol, the number of carbon atoms in the carboxylic acid formed is the same as that of the alcohol. Thus, butanol (four carbon atoms) produces butanoic acid on oxidation.

Answer

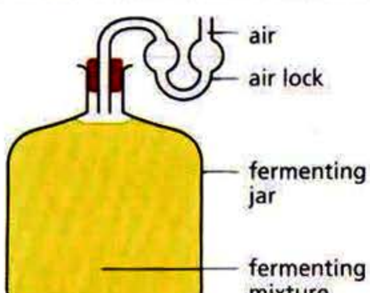


Questions

1. If an alcohol has eight hydrogen atoms,
   a) how many carbon atoms does it have?
   b) what is the name of this alcohol?
   c) how many moles of carbon dioxide will be produced when one mole of this alcohol is burnt completely?

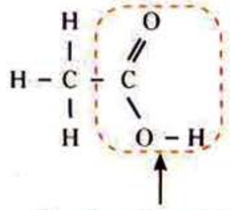
2. An alcohol has the molecular formula $C_4H_{10}O$.
   a) Draw the full structural formula of this alcohol.
   b) Give one physical property and one chemical property of this alcohol.

3. The diagram here shows the apparatus used for fermentation.



   a) Name the substances present in the fermenting mixture.
   b) Write the word equation for the reaction that takes place inside the apparatus.
   c) Describe what will happen if the resulting mixture is exposed to air.

4. State **two** differences between the hydroxide group in sodium hydroxide and the hydroxyl group in alcohols.



All carboxylic acids have this functional group.

Fig. 24.5 The funtional group, -COOH, in carboxylic acids

24.4 | Structure of Carboxylic Acids

What are carboxylic acids?

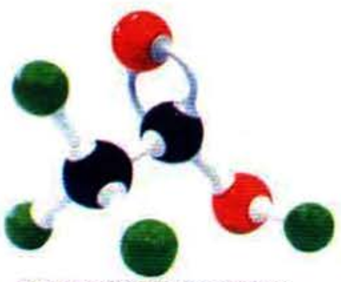
The carboxylic acids are a homologous series of organic compounds with the carboxyl functional group as shown in Fig 24.5.

For simplicity, the functional group is often written as -COOH. Carboxylic acids are weak acids. They are a class of **organic acids**.

Naming Carboxylic Acids

The name of a carboxylic acid ends with '-oic acid'. Most of us are familiar with **vinegar**. Vinegar contains ethanoic acid, which is a carboxylic acid with two carbon atoms.

Table 24.4 shows the formulae of the first four carboxylic acids.



A 3-D model of ethanoic acid

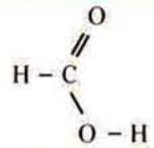
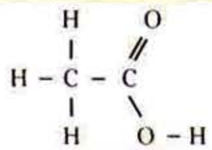
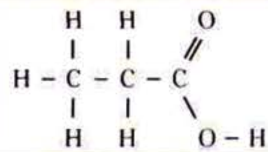
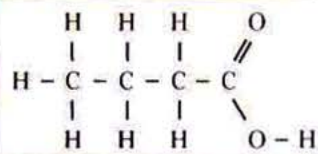
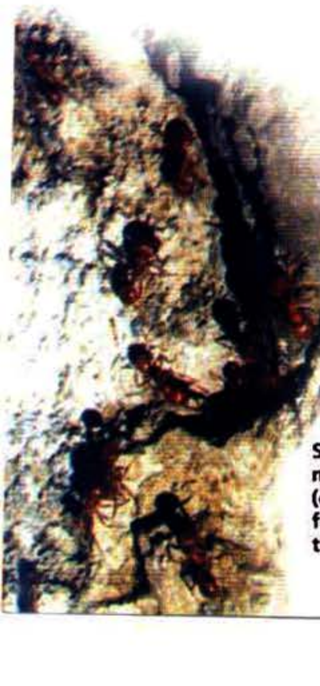
Name	Molecular formula	Structural formula	Full structural formula
Methanoic acid	CH_2O_2	HCOOH	
Ethanoic acid	$C_2H_4O_2$	CH_3COOH	
Propanoic acid	$C_3H_6O_2$ or C_2H_5COOH	CH_3CH_2COOH	
Butanoic acid	$C_4H_8O_2$ or C_3H_7COOH	$CH_3CH_2CH_2COOH$	

Table 24.4 The first four carboxylic acids

What is the general formula for carboxylic acids?

From the structural formula of the carboxylic acids given in Table 24.4, we can conclude that the general formula of carboxylic acids is written as $\mathbf{C_nH_{2n+1}COOH}$. Using this general formula, we can therefore predict that a carboxylic acid with six carbon atoms has the molecular formula $C_5H_{(2\times5)+1}COOH$, that is, $C_5H_{11}COOH$.



Some ants spray methanoic acid (commonly called formic acid) to sting their attackers.

Producing Ethanoic Acid

Ethanoic acid is formed by the oxidation of ethanol. There are two ways to produce ethanoic acid.

1. Ethanoic acid is produced when ethanol is oxidised by atmospheric oxygen in the presence of certain bacteria.
2. In the laboratory, ethanoic acid is prepared by heating a mixture of ethanol and acidified potassium dichromate(VI) (see section 24.2).

A dog can be trained to track people because its sensitive nose can recognise the characteristic smell of human sweat. Each person's sweat contains a unique blend of carboxylic acids.

24.5 | Properties of Carboxylic Acids

Properties of Ethanoic Acid

Ethanoic acid (CH_3COOH) is a colourless liquid at room temperature. It has a characteristic vinegar smell.

Like ethanol, ethanoic acid is completely miscible with (soluble in) water. An aqueous solution of ethanoic acid is a weak acid.

Physical Properties of Carboxylic Acids

Table 24.5 shows the boiling points and solubilities of some carboxylic acids.

Acid	Boiling point (°C)	Solubility in water
Methanoic acid	101	very soluble
Ethanoic acid	118	very soluble
Propanoic acid	141	very soluble
Butanoic acid	164	very soluble

Table 24.5 The boiling points and solubilities of some carboxylic acids

Carboxylic acids that contain more carbon atoms have higher boiling points.

Chemical Properties of Carboxylic Acids

The functional group of carboxylic acid is the carboxyl group, $-\overset{\overset{\displaystyle O}{\|}}{C}-O-H$, or more simply, $-COOH$. It is the carboxyl group that is reactive.

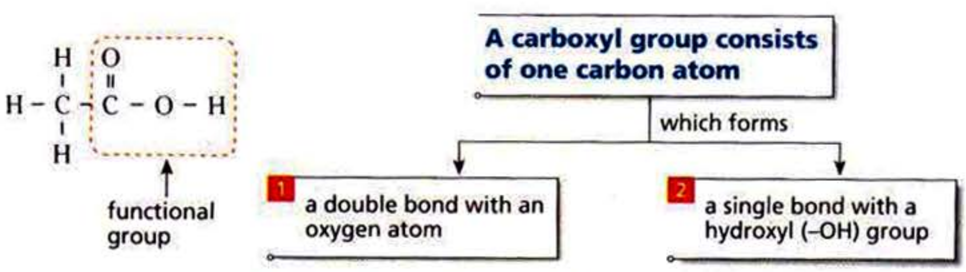


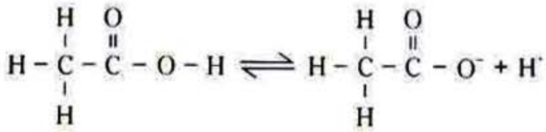
Fig. 24.6 The carboxyl group in ethanoic acid

Quick Check

Both CH_3CH_2OH and CH_3COOH contain –OH group, but only CH_3COOH conducts electricity in aqueous solution. Why?

Carboxylic acids, for example, ethanoic acid, are weak acids because they dissociate only partially in water to form hydrogen ions.

ethanoic acid $\rightleftharpoons$ ethanoate ion + hydrogen ion



It is the hydrogen ions that give carboxylic acids their acidic properties.

1. Reactions with metals

Ethanoic acid reacts with reactive metals such as sodium, potassium and magnesium to give a salt and hydrogen. The salts of ethanoic acid are known as **ethanoates**.

ethanoic acid + magnesium → magnesium ethanoate + hydrogen

$$2CH_3COOH(aq) + Mg(s) \longrightarrow (CH_3COO)_2Mg(aq) + H_2(g)$$

2. Reactions with carbonates

Ethanoic acid reacts with carbonates to give a salt, carbon dioxide and water.

ethanoic acid + sodium carbonate → sodium ethanoate + carbon dioxide + water

$$2CH_3COOH(aq) + Na_2CO_3(aq) \longrightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O(l)$$

3. Reactions with bases

Ethanoic acid reacts with bases to form a salt and water.

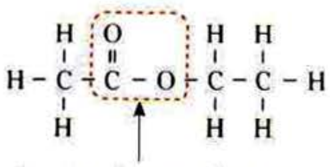
ethanoic acid + copper(II) oxide → copper(II) ethanoate + water

$$2CH_3COOH(aq) + CuO(s) \longrightarrow (CH_3COO)_2Cu(aq) + H_2O(l)$$

24.6 | Esters

What is an ester?

A carboxylic acid reacts with an alcohol to form an organic compound called an ester. This reaction is known as **esterification**.



functional group of an ester

During the formation of an ester, the –OH group of the carboxylic acid is replaced by the CH_3–O– (or C_2H_5–O–, etc.) group of the alcohol.

How do we name an ester?

The name of an ester consists of two parts. For example,

ethyl ethanoate

The first part tells us that the alcohol, ethanol, was used to prepare the ester.	The second part tells us that the carboxylic acid, ethanoic acid, was used to prepare the ester.

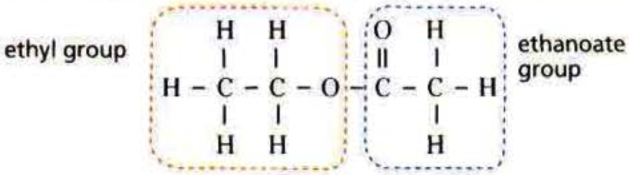


Table 24.6 shows the names of some common esters, together with the alcohol and carboxylic acid used to prepare them.

Name of ester	Structural formula	Alcohol used	Acid used
Ethyl ethanoate	$CH_3COOC_2H_5$	Ethanol	Ethanoic acid
Methyl ethanoate	CH_3COOCH_3	Methanol	Ethanoic acid
Ethyl methanoate	$HCOOC_2H_5$	Ethanol	Methanoic acid

Table 24.6 Some common esters

Uses of Esters

Esters are a very important group of organic compounds. They are found in vegetable oils, animal fats and natural waxes. Esters are used extensively as solvents for cosmetics and glues.

Esters are colourless, neutral liquids which are insoluble in water. They have a sweet, fruity smell. The sweet smell of fruits and flowers is due to the esters present in them. Hence, esters are used in the preparation of perfumes and artificial food flavourings.

Animal fats and vegetable oils are naturally occuring esters. When fats or oils are boiled with sodium hydroxide, the ester breaks up to form glycerol (alcohol) and soap (sodium salt of long-chain carboxylic acid).

TidBit

Rancid butter has a foul smell because of the presence of butanoic acid. However, the esters of butanoic acid have the fragrant smell of fruits. For example, methyl butanoate smells like apples and ethyl butanoate smells like pineapples.

Key ideas

1. Carboxylic acids have the general formula $C_nH_{2n+1}COOH$. They behave as weak acids, reacting with reactive metals, carbonates and bases.
2. Ethanoic acid is the second member of the carboxylic acids. It has the molecular formula CH_3COOH. Ethanoic acid is formed by the oxidation of ethanol. It can be oxidised to ethanoic acid by
   a) bacteria in the air,
   b) hot acidified potassium dichromate(VI) solution.
3. Esters are used in perfumes, solvents and food flavourings.

Test Yourself 24.2

Worked Example

Compound *A* (C_2H_6O) is an alcohol. On oxidation, *A* produces *B*. *B* reacts with *A* to form *C* ($C_4H_8O_2$). *C* is a sweet smelling liquid. The reaction scheme as shown below summarises the reactions.

$$A\ (C_2H_6O) \xrightarrow{\text{oxidising agent}} B \xrightarrow{+A\ (C_2H_6O)} C\ (C_4H_8O_2)$$

a) Identify *A*, *B* and *C*.
b) Write the equations for the reactions.

Thought Process

The molecular formula of *A* shows that it is ethanol. Oxidation of an alcohol produces a carboxylic acid. Hence, *B* is ethanoic acid.
A sweet-smelling liquid is most likely an ester. *C* is produced from *A* and *B*, so *C* is most likely ethyl ethanoate.

Answer

a) *A* is ethanol, CH_3CH_2OH.
B is ethanoic acid, CH_3COOH.
C is ethyl ethanoate, $CH_3COOC_2H_5$.

b) $CH_3CH_2OH + 2[O] \longrightarrow CH_3COOH + H_2O$

$CH_3COOH + CH_3CH_2OH \longrightarrow CH_3COOC_2H_5 + H_2O$

Questions

1. Describe and explain what happens when a test tube containing ethanol and two drops of blue litmus solution are exposed to the air for about three days.

2. a) Give the conditions necessary for the following conversions:

$$CH_2 = CH_2 \xrightarrow{\text{I}} C_2H_5OH \xrightarrow{\text{II}} CH_3CO_2C_2H_5$$

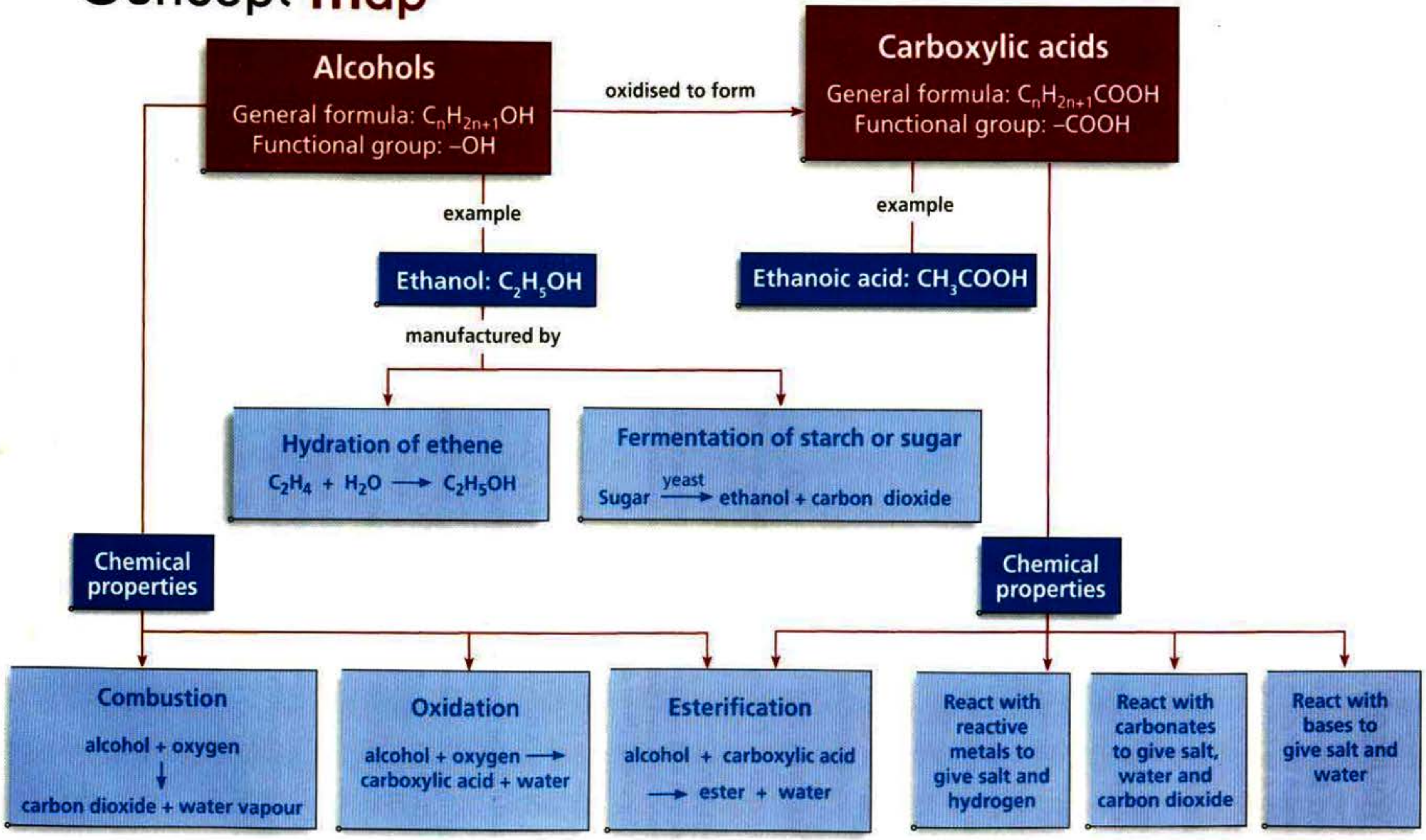
b) Name the processes involved in I and II.

3. Identify substances *A*, *B*, *C* and *D* from the following reactions:
a) $NaOH + A \longrightarrow HCOONa + H_2O$
b) $CH_3OH + B \longrightarrow CH_3COOCH_3 + H_2O$
c) $2CH_3CH_2COOH + Mg \longrightarrow C\ (gas) + D$

4. Esters have fruity smells and are often manufactured. For each of the following esters, write the structural formulae of the organic compounds that are needed to prepare the ester.

a) $CH_3COOCH_2CH_2CH_3$ which smells like pears
b) $CH_3COOCH_2CH_2CH_2CH_2CH_3$ which smells like bananas

Concept Map


Alcohols
General formula: $C_nH_{2n+1}OH$
Functional group: –OH
oxidised to form
Carboxylic acids
General formula: $C_nH_{2n+1}COOH$
Functional group: –COOH
example
Ethanol: C_2H_5OH
example
Ethanoic acid: CH_3COOH
manufactured by
Hydration of ethene
$C_2H_4 + H_2O \longrightarrow C_2H_5OH$
Fermentation of starch or sugar
Sugar $\xrightarrow{\text{yeast}}$ ethanol + carbon dioxide
Chemical properties
Chemical properties
Combustion
alcohol + oxygen $\downarrow$ carbon dioxide + water vapour
Oxidation
alcohol + oxygen $\longrightarrow$ carboxylic acid + water
Esterification
alcohol + carboxylic acid $\longrightarrow$ ester + water
React with reactive metals to give salt and hydrogen
React with carbonates to give salt, water and carbon dioxide
React with bases to give salt and water

Exercise 24

Foundation

1. Ethanol and ethene have the same ________.

   A mass B number of carbon atoms
   C volume D number of molecules

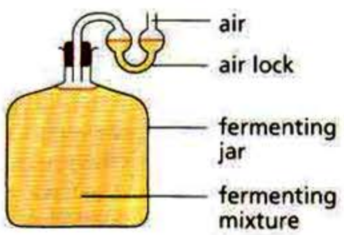
2. Identify the equation which represents the complete combustion of C_3H_7OH.

   A $C_3H_7OH + O_2 \longrightarrow C_2H_5CO_2H + H_2O$
   B $2C_3H_7OH + 5O_2 \longrightarrow 6CO_2 + 8H_2$
   C $2C_3H_7OH + 6O_2 \longrightarrow 6CO + 8H_2O$
   D $2C_3H_7OH + 9O_2 \longrightarrow 6CO_2 + 8H_2O$

3. Which of the following statements about the production of ethanol is **not** true?

   A Ethanol can be made by the fermentation of starch.
   B Oxygen is evolved during fermentation.
   C Fermentation requires yeast.
   D Fermentation proceeds best at about 37 °C.

4. The apparatus below is used for fermentation.



   The function of the air lock is to stop substance *X* from getting into the jar and to let substance *Y* out. What are *X* and *Y*?

	X	*Y*
A	carbon dioxide	oxygen
B	ethanol	carbon dioxide
C	nitrogen	carbon dioxide
D	oxygen	carbon dioxide

5. When a mixture of ethanol and acidified potassium dichromate(VI) is heated strongly, the colour of the solution changes from ________.

   A orange to green
   B green to orange
   C brown to colourless
   D colourless to brown

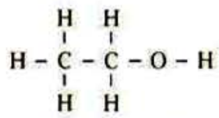
6. On exposure to air, the taste of wine changes because the alcohol in the wine ________.

   A is oxidised by oxygen in the air
   B is reduced by oxygen in the air
   C is oxidised by carbon dioxide in the air
   D is reduced by carbon dioxide in the air

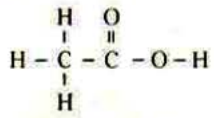
7. When ethanol is boiled with acidified potassium dichromate(VI), it produces ________.

   A carbon monoxide B ethane
   C ethanoic acid D ethene

8. The full structural formulae of compounds *X* and *Y* are shown below.



Compound *X*

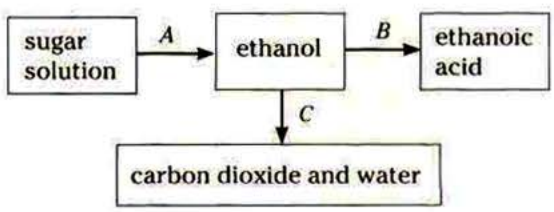


Compound *Y*

   The best method to distinguish between *X* and *Y* is by using ________.

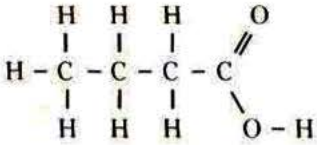
   A bromine water
   B sodium hydroxide solution
   C sodium carbonate solution
   D dilute sulphuric acid

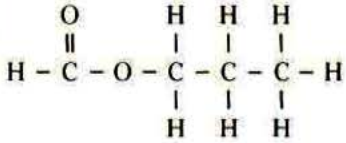
9. a) Name reactions *A*, *B* and *C* as shown in the reaction scheme.

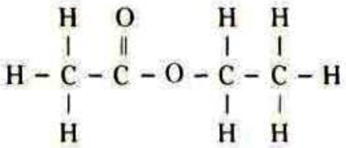


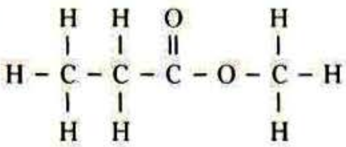
   b) Draw the structural formula of
      i) ethanol.
      ii) ethanoic acid.

10. Which is the structural formula of ethyl ethanoate?

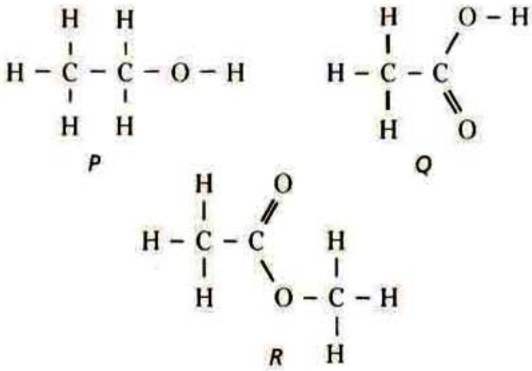
A


B


C


D


11. The structures of three organic compounds, *P*, *Q* and *R* are shown below.



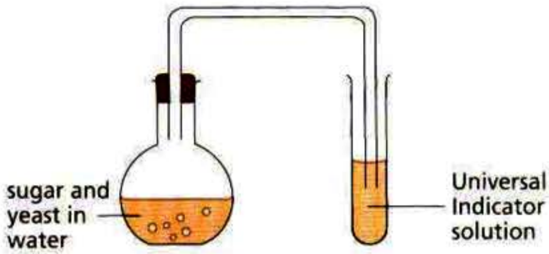
a) Describe an industrial process to make *P* from ethene.
b) Name (i) the reagents, (ii) the condition that can be used to convert *P* to *Q*.
c) Name the type of reaction that occurs when *P* is converted to *Q*.
d) Name (i) the reagent, (ii) the catalyst and (iii) the conditions for the conversion of *Q* to *R*.
e) Name the type of reaction that occurs when *Q* is converted to *R*.

Challenge

1. Ethanol is used as a fuel for cars in some countries that grow a lot of sugar cane. First, the sugar cane is crushed and mixed with water to dissolve the sugar. Yeast is then added to the concentrated sugar solution. When the reaction stops, the yeast is removed. A solution of ethanol and water is left behind.

   a) What is the name of this method of making ethanol from sugar solution?
   b) Draw a labelled diagram of how you would remove the yeast and collect the solution of ethanol.
   c) Explain why it is possible to separate ethanol from the solution by distillation after the yeast is removed.

2. A student prepares a solution of ethanol as shown below.



   a) i) What name is given to this process by which sugars are converted to ethanol?
      ii) The student wants the process to go faster, so he heats the flask. Explain why this may not be a good idea.
   b) A gas that is produced is bubbled through the Universal Indicator solution, turning it from green (pH 7) to orange (pH 4).
      i) What does this tell you about the gas that is produced?
      ii) Suggest one other function of the test tube containing the indicator solution.

3. When fats and oils are heated with concentrated sodium hydroxide, glycerol is produced. The formula of glycerol is:

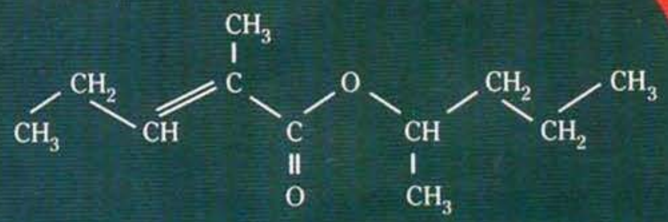
   $CH_2OHCHOHCH_2OH$

   a) Draw the structure of glycerol.
   b) In what way is the structure of glycerol different from alcohols such as methanol and ethanol?

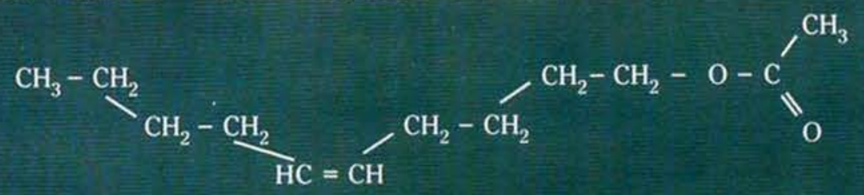
Chemistry Today

Pheromones are organic compounds emitted by insects and other animals to send messages to individuals of the same species. Researchers have identified and isolated pheromones from silkworm moths, butterflies, bees and some beetles.

Insects use pheromones to attract mates, to signal danger, to mark their territory and to lay trails that other insects can follow. For example, female moths release pheromones that are picked up by the feathery antennae of male moths. Male moths can follow a pheromone trail to find a female moth several miles away!



Pheromone released by beetles



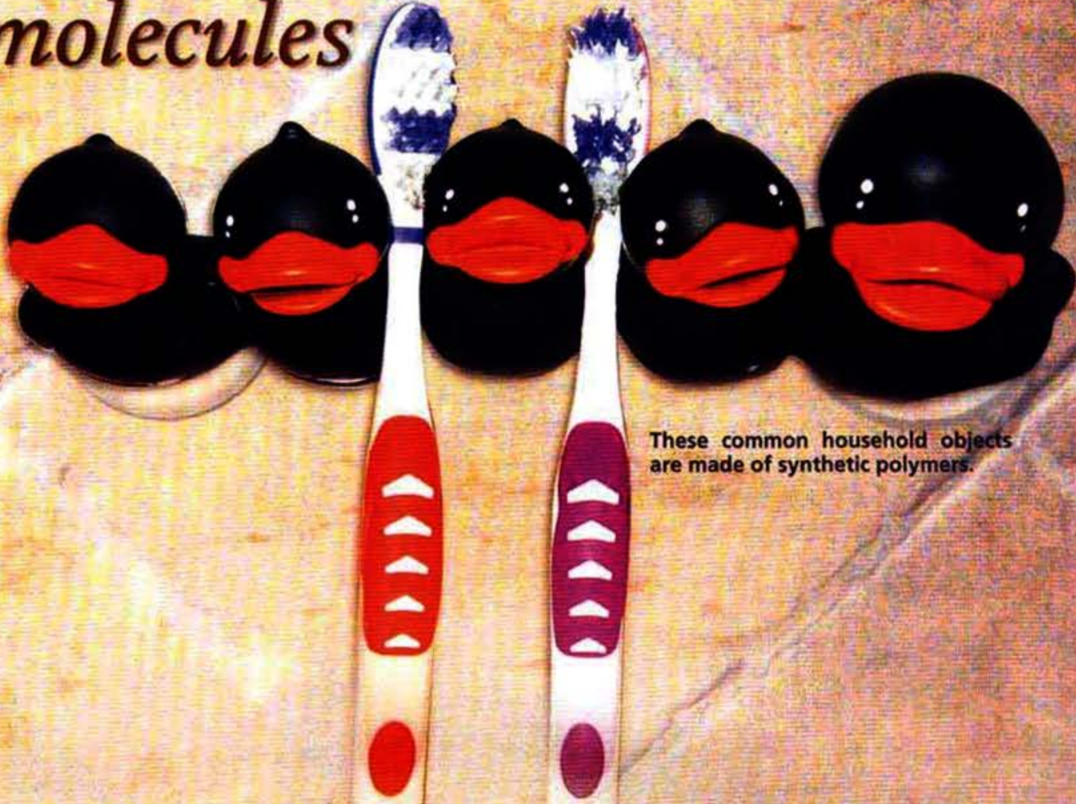
Pheromone released by butterflies

CRITICAL THINKING

a) Based on the formulae of the pheromones given, state three general conclusions about insect pheromones.
b) State one physical property that a pheromone must have, in order for it to function as a danger signal or to attract mates. (Hint: How do the pheromone 'messages' reach other individuals?)
c) Pheromones can be used for pest control. Suggest how this can be carried out.

Chapter 25

Macromolecules



These common household objects are made of synthetic polymers.

Most of the organic compounds that we studied in the earlier chapters are composed of simple molecules. These simple molecules contain only a few atoms. However, there exist complex molecules which contain thousands or even millions of atoms. These molecules are called **macromolecules**.

Macromolecules are all around us. Examples of macromolecules are starch and proteins, which are necessary for life processes. These macromolecules occur naturally.

Since the last century, chemists have been able to create macromolecules with specific properties to meet our needs. This new class of compounds is known as **synthetic polymers**. They include plastics such as poly(ethene), nylon and Terylene.

Chapter Outline

25.1 | Macromolecules

Our health and lifestyle depend to a large extent on a class of organic compounds known as polymers. The food we eat, the clothes we wear and many of the things we use are made of very large molecules called polymers. Polymers are macromolecules.

What is a macromolecule?

A **macromolecule** is any long-chain molecule that contains hundreds or thousands of atoms, joined together by covalent bonds. A macromolecule is formed by linking together many small repeating units known as **monomers**. *The process of joining together a large number of small molecules (monomers) to form a macromolecule* is called **polymerisation**. The macromolecule formed is called a **polymer**.

Chem-Aid

'Poly' means 'many'; 'mer' means 'unit'; 'mono' means 'one.

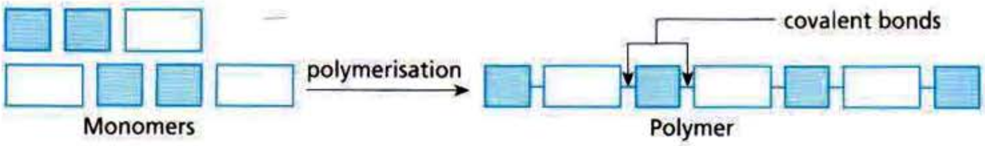


Fig. 25.1 Polymerisation

Natural and Synthetic Polymers

The first polymers to be used were those which occur naturally. For example, the natural fibres produced by silkworms have been used for thousands of years. However, in recent years, advances in science have enabled scientists to produce a wide range of man-made polymers. Table 25.1 shows some examples of natural and synthetic (man-made) polymers.

Natural polymers	Synthetic polymers
• proteins • carbohydrates • fats	• poly(ethene) • nylon • Terylene • polyvinyl chloride (PVC)

Table 25.1 Examples of natural and synthetic polymers

25.2 | Addition Polymerisation

There are two basic types of reactions for forming polymers: addition polymerisation and condensation polymerisation. **Addition polymerisation** *occurs when monomer units join together without losing any molecules or atoms.* Alkenes can undergo addition polymerisation. At high pressure (1000 atm) and high temperature (200 °C), thousands of alkene molecules join together to form polymers.

Poly(ethene)

The simplest addition polymer is **poly(ethene)**. Poly(ethene) is commonly known as **polythene**. It is the material used for making bags, toys, buckets, clingfilm etc. It is flexible but difficult to break. Poly(ethene) is produced by addition polymerisation.

Link

In addition polymerisation, an alkene molecule reacts with other alkene molecules. What other substances react with alkenes by addition reactions? See chapter 23.

TidBit

Different types of poly(ethene) can be produced by using different conditions.

- HDPE – high density poly(ethene)
- MDPE – medium density poly(ethene)
- LDPE – low density poly(ethene)
- LLDPE – linear low density poly(ethene)

A 3-D model showing a section of poly(ethene)

Karl Ziegler (1898 – 1973)

Ziegler's most important discovery was made in 1953, when he found a catalyst for the polymerisation of alkenes. The catalyst is now called Ziegler's catalyst. It is a mixture of triethyl aluminium, $(C_2H_5)_3Al$, and titanium tetrachloride, $TiCl_4$. Using Ziegler's catalyst, the polymerisation of ethene can be carried out at a much lower temperature and pressure, and high density polythene (HDPE) is obtained.

What happens in addition polymerisation to form poly(ethene)?

At high temperature and pressure, and in the presence of a catalyst, the carbon-carbon double bonds of the alkene molecules (monomers) break. Each monomer forms single bonds with two other monomers. Eventually, they join to form a giant molecule called a polymer.

This is what happens during the addition polymerisation of ethene to form the polymer, poly(ethene).

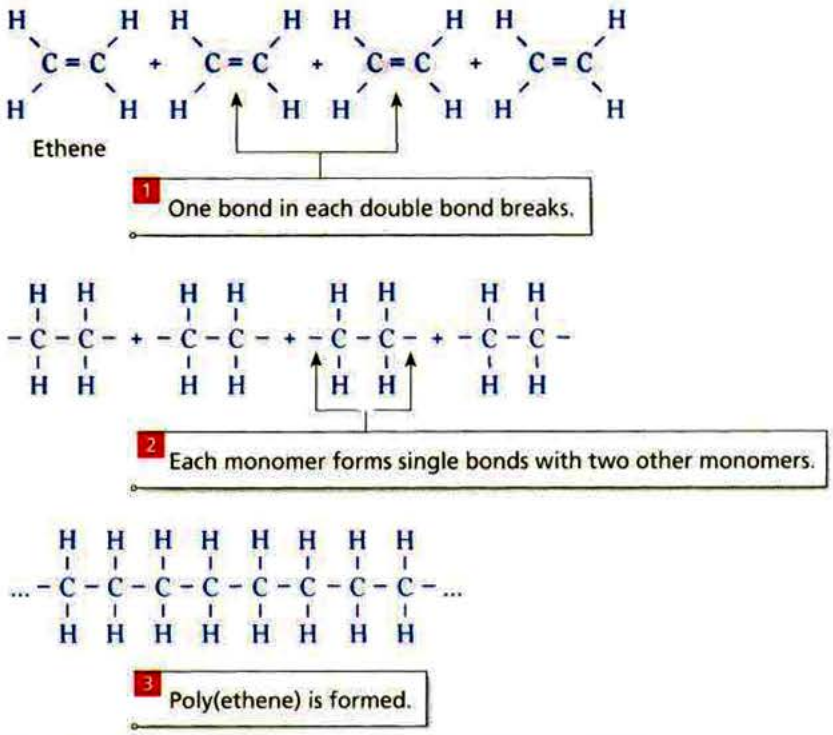


Fig. 25.2 Addition polymerisation of ethene to form poly(ethene)

The equation for the polymerisation of ethene is therefore written as

$$\text{ethene} \xrightarrow[\text{catalyst}]{\text{high temperature and pressure}} \text{poly(ethene)}$$

OR

$$(n\,CH_2{=}CH_2) \xrightarrow[\text{catalyst}]{\text{high temperature and pressure}} \text{-(}CH_2{-}CH_2\text{)-}_n$$

OR

$$n\begin{array}{c}H\\|\\C\\|\\H\end{array}=\begin{array}{c}H\\|\\C\\|\\H\end{array} \xrightarrow[\text{catalyst}]{\text{high temperature and pressure}} \left[\begin{array}{c}H\\|\\C\\|\\H\end{array}-\begin{array}{c}H\\|\\C\\|\\H\end{array}\right]_n$$

The repeat unit of poly(ethene) is $\left[\begin{array}{c} \mathrm{H} \quad \mathrm{H} \\ | \quad\; | \\ -\mathrm{C}-\mathrm{C}- \\ | \quad\; | \\ \mathrm{H} \quad \mathrm{H} \end{array}\right]_n$. The formula of a polymer can be written using its repeat unit. Thus, poly(ethene) can be written simply as $\left[\begin{array}{c} \mathrm{H} \quad \mathrm{H} \\ | \quad\; | \\ -\mathrm{C}-\mathrm{C}- \\ | \quad\; | \\ \mathrm{H} \quad \mathrm{H} \end{array}\right]_n$ or $(CH_2-CH_2)_n$.

The letter 'n' represents a large number. The number of ethene molecules in each poly(ethene) molecule varies between 10 000 and 30 000.

Poly(ethene) is a type of material known as plastic. You will learn more about plastics in section 25.4.

How do we deduce the structural formula of the monomer from its polymer?

In the previous section, you learnt how to draw a polymer given its monomer. How can we do the reverse? Let's examine this using PVA (polyvinyl acetate) as an example. PVA is used as the main constituent in emulsion paints. It has the formula

$$\begin{array}{llllll} & \mathrm{OCOCH_3} & & \mathrm{OCOCH_3} & & \mathrm{OCOCH_3} \\ & \;\;| & & \;\;| & & \;\;| \\ \ldots - & \mathrm{CH} \;—— & \mathrm{CH_2} - & \mathrm{CH} \;—— & \mathrm{CH_2} - & \mathrm{CH} \;—— \ldots \end{array}$$

To deduce the structural formula of its monomer, we follow the steps below.

1. Identify the repeat unit in the polymer.

$$\begin{array}{llllll} & \mathrm{OCOCH_3} & & \mathrm{OCOCH_3} & & \mathrm{OCOCH_3} \\ & \;\;| & & \;\;| & & \;\;| \\ \ldots - & \mathrm{CH} \;—— & \mathrm{CH_2} - & \mathrm{CH} \;—— & \mathrm{CH_2} - & \mathrm{CH} \;—— \ldots \end{array}$$

|← Repeat unit →| |← Repeat unit →|

2. Write down the formula of the repeat unit.

$$\begin{array}{cc} \mathrm{H} & \mathrm{OCOCH_3} \\ | & | \\ -\mathrm{C} - & \mathrm{C} - \\ | & | \\ \mathrm{H} & \mathrm{H} \end{array}$$

3. Convert the carbon-carbon single bond into a carbon-carbon double bond. The structural formula of the monomer, vinyl acetate, is

$$\begin{array}{cc} \mathrm{H} & \mathrm{OCOCH_3} \\ | & | \\ \mathrm{C} = & \mathrm{C} \\ | & | \\ \mathrm{H} & \mathrm{H} \end{array}$$

Vinyl acetate

Quick Check

1. **State two differences between ethene and poly(ethene).**
2. **Write down the formula of the polymer obtained from the addition polymerisation of 1,1-dichloroethene, $CH_2=CCl_2$.**

Chem-Aid

A plastic is a macromolecule that can be made into different shapes.

The Uses of Addition Polymers

Addition polymers are good insulators of heat and electricity, and are resistant to chemical attack. Some addition polymers and their uses are shown here.

Polystyrene is hard, light and brittle. It is used to make disposable containers.

The windscreens of cars are made of Perspex, a polymer. Perspex is transparent and less easy to break than glass.

Many frying pans are coated with a layer of Teflon which is also known as poly(tetrafluoroethene). Teflon is heat-resistant. It has 'non-stick' properties.

Polyvinyl chloride (PVC) is used to make pipes, raincoats, thin gloves and flooring mats.

Key ideas

1. Macromolecules are long-chain molecules made up of large numbers of small molecules called monomers joined together by covalent bonds.
2. Alkenes can undergo addition polymerisation to form large molecules called addition polymers. Addition polymers are macromolecules.
3. Poly(ethene), polyvinyl chloride (PVC), polystyrene, Perspex and Teflon are examples of addition polymers. They are chemically inert and do not conduct heat and electricity.

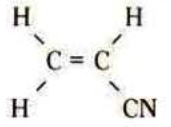
TidBit

Synthetic polymers can be classified into two main groups: thermoplastics and thermosetting plastics. All addition polymers are **thermoplastics**. This means that they melt on heating and can be remoulded into different shapes. **Thermosetting plastics** do not melt on heating. Instead they decompose when heated. Examples of thermosetting plastics include Bakelite and poly(urethane). They are not addition polymers. Bakelite is used for making electric plugs and poly(urethane) is used for making mattresses and cushion pillows.

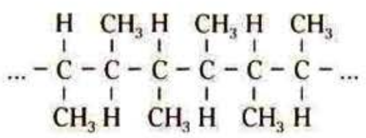
Test Yourself 25.1

Questions

1. Draw the polymer formed by the following monomer.



2. The structure below represents part of a polymer chain. Draw the full structural formula of its monomer.



25.3 | Condensation Polymerisation

Some polymers are made by reacting two different types of monomers. Each of the monomers involved has a functional group at each end of the molecule. When these monomers react, *a polymer is produced, and a small molecule, such as water, is also produced as a by-product of the reaction*. This type of reaction is called **condensation polymerisation**.

There are two main groups of condensation polymers: the **polyamides** and the **polyesters**.

Nylon — A Synthetic Polyamide

Nylon was the first synthetic fibre made by condensation polymerisation. It is made from the monomers dicarboxylic acid and diamine. We can represent these two monomer units as follows:

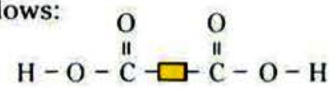


Fig. 25.3 Dicarboxylic acid — a molecule with two -COOH groups

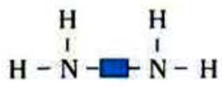


Fig. 25.3 Diamine — a molecule with two $-NH_2$ groups

This is how the two monomers react:

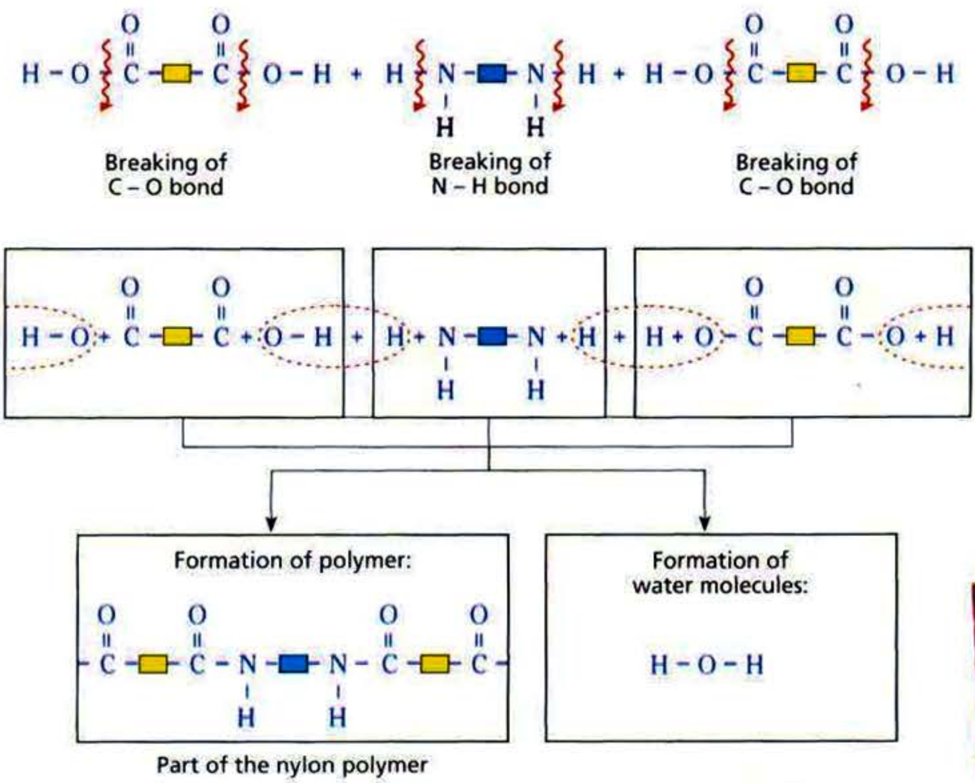


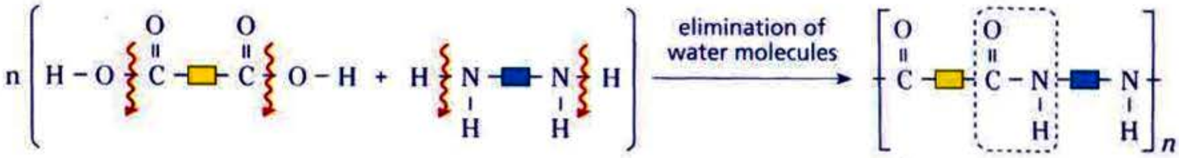
Fig. 25.5 Condensation polymerisation of dicarboxylic acid and diamine to form nylon

Stephanie Kwolek (1923 –)

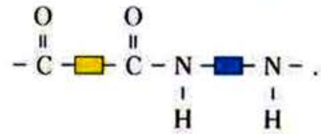
Stephanie Kwolek, an American chemist, spearheaded polymer research at DuPont's Pioneering Research Laboratory for 40 years as a Research Associate. She retired in 1986. Kwolek is best known for her work on the development of Kevlar fibre, a polyamide which is five times stronger than steel. You may learn more about Kevlar in Chemistry Today, page 445.

Nylon is used to make sewing thread because it is strong and can be drawn into long thin strands without breaking.

The reaction between a dicarboxylic acid and a diamine can be represented by the equation shown below.



Nylon

The repeat unit of nylon is 

Notice the following:

1. Each monomer for making nylon has two identical functional groups. The chain can be extended in two directions by the removal of water molecules.
2. A whole family of nylons can be formed using different dicarboxylic acid and diamine monomers.

Quick Check

Nylon can also be prepared by the condensation polymerisation between the monomers

Predict the by-product of this reaction.

The reactive group (or functional group) in nylon is shown below.

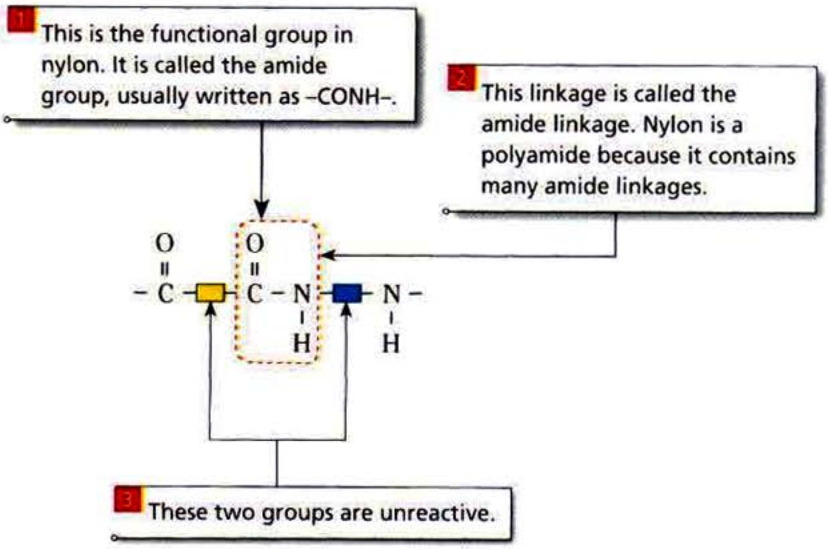


Fig. 25.6 The functional group in nylon

Chem-Aid

Hair, wool, silk and proteins are natural polymers that are polyamides.

TidBit

Every part of your body contains proteins. Enzymes (biological catalysts) are proteins too. Proteins are naturally occurring polyamides. They are made by linking amino acid monomers together. Every amino acid contains an amine group ($-NH_2$) and a carboxylic acid group ($-COOH$).

Terylene — A Synthetic Polyester

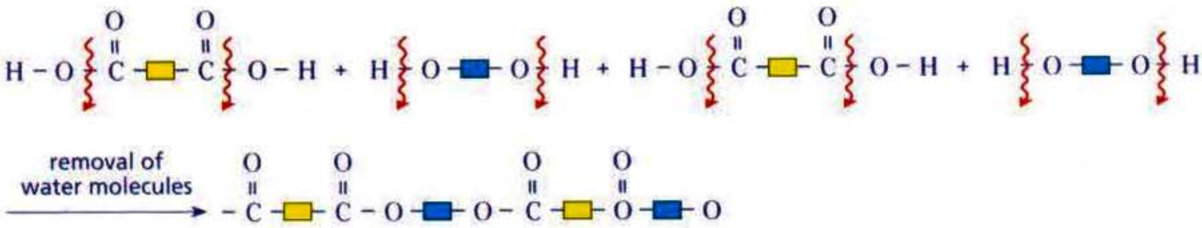
Besides nylon, Terylene is another example of a condensation polymer. Terylene is made from these monomers: benzene–1, 4–dicarboxylic acid and ethane–1, 2–diol. We can represent the acid molecule and the alcohol molecule as follows:

Fig. 25.7 Dicarboxylic acid — an acid with two -COOH groups

Fig. 25.8 Diol — an alcohol with two -OH groups

Fibres of Terylene (also called Dacron) are widely used in clothing and fabrics. Terylene is often mixed with cotton to produce polyester-cotton fabrics.

These monomers react to form Terylene and water:



Part of the Terylene polymer

The reactive group or functional group in Terylene is shown below.

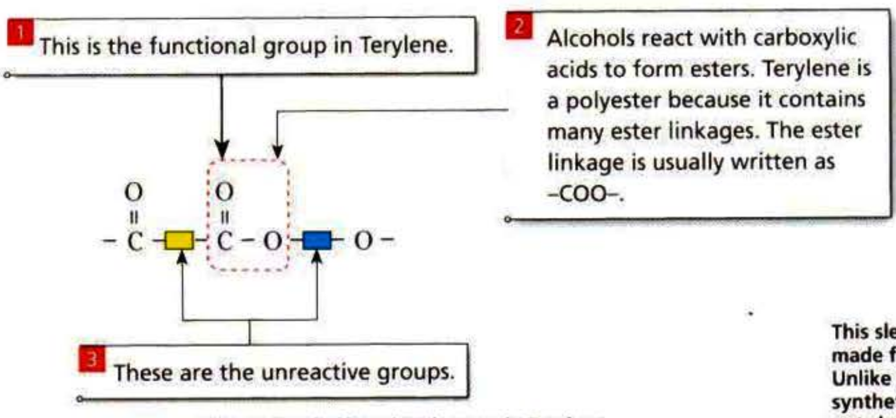
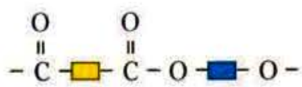


Fig. 25.9 The functional group in Terylene

The repeat unit of Terylene is



This sleeping bag is made from Terylene. Unlike natural fibres, synthetic fibres do not shrink. What are the other advantages of using Terylene to make sleeping bags?

Uses of Man-made Fibres

Synthetic or man-made fibres are usually plastics which have been spun into threads. Terylene and nylon are the best known examples of synthetic fibres. Clothes made from these fibres are shrink-proof and crease-proof. They are also easier to wash and dry.

Examples of items made from nylon and Terylene are curtains, parachutes, fishing lines and sleeping bags.

Try it Out

We can dispose of waste materials by incineration, recycling or landfilling. How are waste materials in Singapore disposed of? Find out from 'Waste Minimisation', at http://app.nea.gov.sg

25.4 | Plastics and Pollution

Addition polymerisation and condensation polymerisation are used to produce plastics such as poly(ethene) and nylon. Plastics are now used in place of natural materials such as wood, metal, cotton and leather because they are

- relatively cheap,
- easily moulded into various shapes,
- light, tough and waterproof,
- durable (resistant to decay, rusting and chemical attack).

What are the disadvantages of using plastics?

Plastics are carbon compounds and hence they are flammable. Modern buildings are insulated and furnished with plastic materials. Unlike wood, metals and ceramics, plastics catch fire easily. *When plastics burn, fires can spread very quickly and poisonous gases are produced.*

Improper disposal of plastic leads to land pollution.

Plastics are **non-biodegradable**. This means that *they cannot be decomposed by bacteria in the soil*. The biggest problem with plastics is: what do we do with them after we have finished using them? We could bury plastic waste in a landfill and get it out of sight. However, plastics do not decompose when they are thrown away. They take up space and *cause land pollution*.

Can we dispose of plastic objects by burning them?

Most plastics can be destroyed by burning (**incineration**). When poly(ethene) is burnt in sufficient air, two non-polluting products are produced: carbon dioxide and water. The heat energy from this burning can be used to supply factories with power.

Try it Out

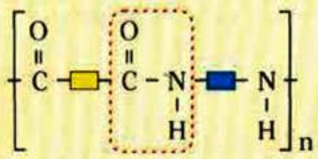
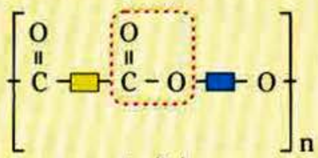
Nylon is a synthetic polymer. It is difficult to decompose nylon and this makes it difficult for us to dispose of it. Can we replace nylon with biodegradable plastics? Use the Internet to find out more about biodegradable plastic.

However, many plastics produce poisonous gases upon incineration. For example, PVC produces hydrogen chloride gas on burning. *Burning plastics will therefore cause air pollution.* The best way to deal with plastic objects is to reuse or recycle them.

Recycling bins are usually placed at convenient locations in schools, offices and public places. This encourages people to recycle materials that are recyclable.

Key ideas

1. In condensation polymerisation, monomer units combine to form a polymer, and small simple molecules such as water are eliminated.
2. Nylon and Terylene are made by condensation polymerisation.
3. Nylon and Terylene are synthetic polymers which are often made into fibres. They are used for making clothes, curtain materials e.t.c.
4. The two disadvantages of plastic are: they are fire hazards and they cause pollution problems due to their non-biodegradable nature.
5. The table below shows some synthetic polymers and their monomers.

Polymer	Formula of polymer and type of linkage	Monomer units
Nylon	 amide linkage	HOOC –▭– COOH (dicarboxylic acid) and H_2N –▭– HN_2 (diamine)
Terylene	 ester linkage	HOOC –▭– COOH (dicarboxylic acid) and HO –▭– OH (diol)

Test Yourself 25.2

Worked Example

There are many types of nylons. The structure of the monomer for making nylon-6 is shown here.

$$H-\underset{}{N}(H)-(CH_2)_5-C(=O)-O-H$$

Draw the repeat unit of nylon-6.

Thought Process

The monomer contains two functional groups: amine group and carboxylic acid group. (This is unlike the example that you have studied.) The monomers will undergo condensation polymerisation.

$$\underbrace{H-N(H)}_{\text{Amine group}}-(CH_2)_5-\underbrace{C(=O)-O-H}_{\text{Carboxylic acid group}}$$

In the condensation polymerisation between amine and carboxylic acid, water molecules are removed. Hence,

$$H-N(H)-(CH_2)_5-C(=O)-O-H + H-N(H)-(CH_2)_5-C(=O)-O-H \longrightarrow \text{nylon-6} + H_2O$$

Answer

The repeat unit of nylon-6 is: $-N(H)-(CH_2)_5-C(=O)-$

Questions

1. Name the elements present in (a) nylon and (b) Terylene.

2. State **two** differences between condensation polymerisation and addition polymerisation.

3. Plastic waste can be disposed of by burning it in an incinerator. Give one advantage and one drawback of this method.

4. A part of the nylon molecule is shown below.

a) Give the name and structure of the linkage in this molecule.
b) Draw the repeat unit of this polymer.

5. Part of the Terylene molecule is shown below.

Write down the name and full structural formula of *X*.

Concept Map

Macromolecules — can cause pollution →

Pollution caused by non-biodegradable plastics:

- non-biodegradable — cannot be decomposed by bacteria in the soil.

↓

Problems:

- fire hazard
- may produce poisonous gases when burnt
- cause land pollution

Macromolecules — are →

Large molecules built up from small units

— two types produced by →

Addition polymerisation:

- occurs when monomer units join together without losing any molecules
- example:
  ethene ⟶ poly(ethene)
- uses: plastic bags, raincoats e.t.c.

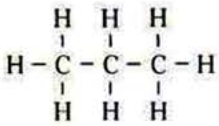
Condensation polymerisation:

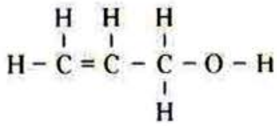
- occurs when monomer units join together with the elimination of simple molecules such as water.
- examples:
  nylon (polyamide) and Terylene (polyester)
- uses: as man-made fibres for making clothing, fishing lines, parachutes e.t.c.

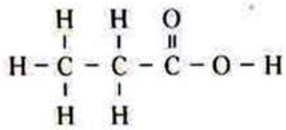
Exercise 25

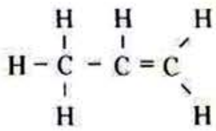
Foundation

1. Which of the following compounds undergo addition polymerisation?

1) 

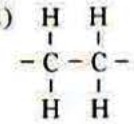
2) 

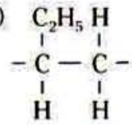
3) 

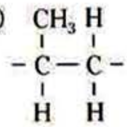
4) 

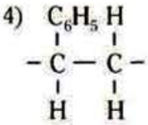
A 1 and 3 only. B 2 and 4 only.
C 1, 2 and 3 only. D 2, 3 and 4 only.

2. Which of the following have the empirical formula CH_2?

1) 

2) 

3) 

4) 

A 1 and 3 only. B 2 and 4 only.
C 1, 2 and 3 only. D 2, 3 and 4 only.

3. The biggest disadvantage of addition polymers is the fact that ____________.

A they have low melting points
B they have limited applications
C they cannot be decomposed by bacteria
D they are easily attacked by acids

4. Which of the following compounds form polymers?

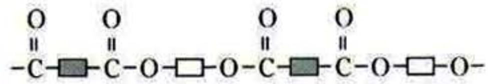
A C_2H_6 B C_3H_3
C C_4H_{10} D C_2H_6O

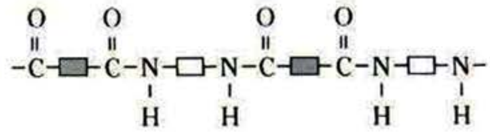
5. Which of the following **cannot** undergo condensation polymerisation?

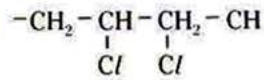
A *Cl*–□–*Cl* B HO–□–OH
C HO–□–COOH D HOOC–□–COOH

6. Which of the following diagrams best represents the structure of Terylene?

A 

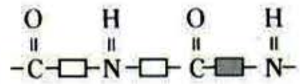
B 

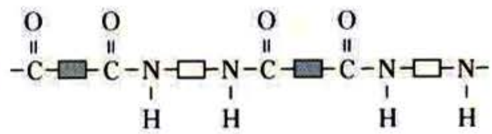
C 

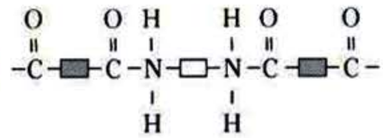
D 

7. Which of the following diagrams best represents the structure of nylon?

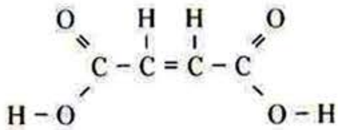
A 

B 

C 

D 

8. The structural formula of butenedioic acid is shown below.



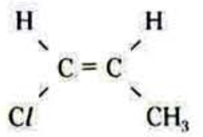
Which statement about butenedioic acid is correct?

A Its empirical formula is the same as its molecular formula.
B It can only undergo addition polymerisation.
C It can only undergo condensation polymerisation.
D It can undergo both addition polymerisation and condensation polymerisation.

9. Which of the following plastics produces a poisonous acidic gas when burnt in the incinerator?

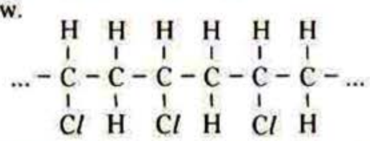
   A Poly(ethene)
   B Poly(phenylethene)
   C Poly(chloroethene)
   D Terylene

10. a) The following structure represents a monomer.



Draw the structure of the polymer which shows **three** repeat units.

b) A polymer has the structure as shown below.



i) How many repeat units are shown in the structure?
ii) What is the molecular formula of the monomer?

11. Nylon is a polymer produced by condensation polymerisation.

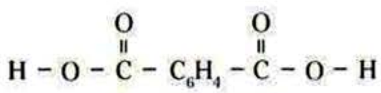
   a) Nylon is a polyamide. What is meant by polyamide?

   b) Write the structure to show a section of the nylon polymer.

12. PET or PETE is the abbreviation for poly(ethylene terephthalate). It is produced by the reaction between ethylene glycol and terephthalic acid.

$H - O - (CH_2)_2 - O - H$

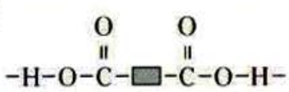
Ethylene glycol



Terephthalic acid

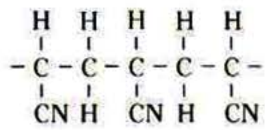
   a) What is the name for this type of reaction?

   b) Write the structural formula of PET. Hint: you may visualise ethylene glycol as H – O –▭– O – H and terephthalic acid as



Challenge

1. Poly(acrylonitrile) is used as a synthetic fibre under the trade name, Orlon. Part of the structure of poly(acrylonitrile) is shown below.



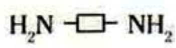
   a) Draw the full structural formula of acrylonitrile.

   b) What type of polymerisation is responsible for its formation?

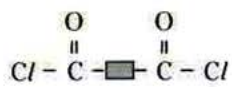
   c) Predict and explain the electrical conductivity of poly(acrylonitrile) in the molten state.

   d) Name **two** gases that are produced when poly(acrylonitrile) is burnt in excess oxygen.

2. Consider two compounds, *X* and *Y*, with the structures as shown below.



X



Y

When *X* and *Y* react, hydrogen chloride gas is liberated and a compound, *Z*, is produced.

   a) State the type of reaction that occurs between *X* and *Y*.

   b) i) Draw the structural formula of compound *Z*.
   ii) State **one** use of compound *Z*.

   c) Name a synthetic fibre that has the same structure as compound *Z*.

Chemistry Today

Most plastics that we encounter are either relatively soft or break easily. In the early 1970s, a very strong polymer, known as Kevlar, was created. Kevlar is produced by reacting an alcohol with a carboxylic acid. Kevlar is made from two different monomers. One of the monomers used to make Kevlar is shown here.

$$HO-\overset{\overset{\displaystyle O}{\|}}{C}-C_6H_4-\overset{\overset{\displaystyle O}{\|}}{C}-OH$$

1,4-benzenedicarboxylic acid

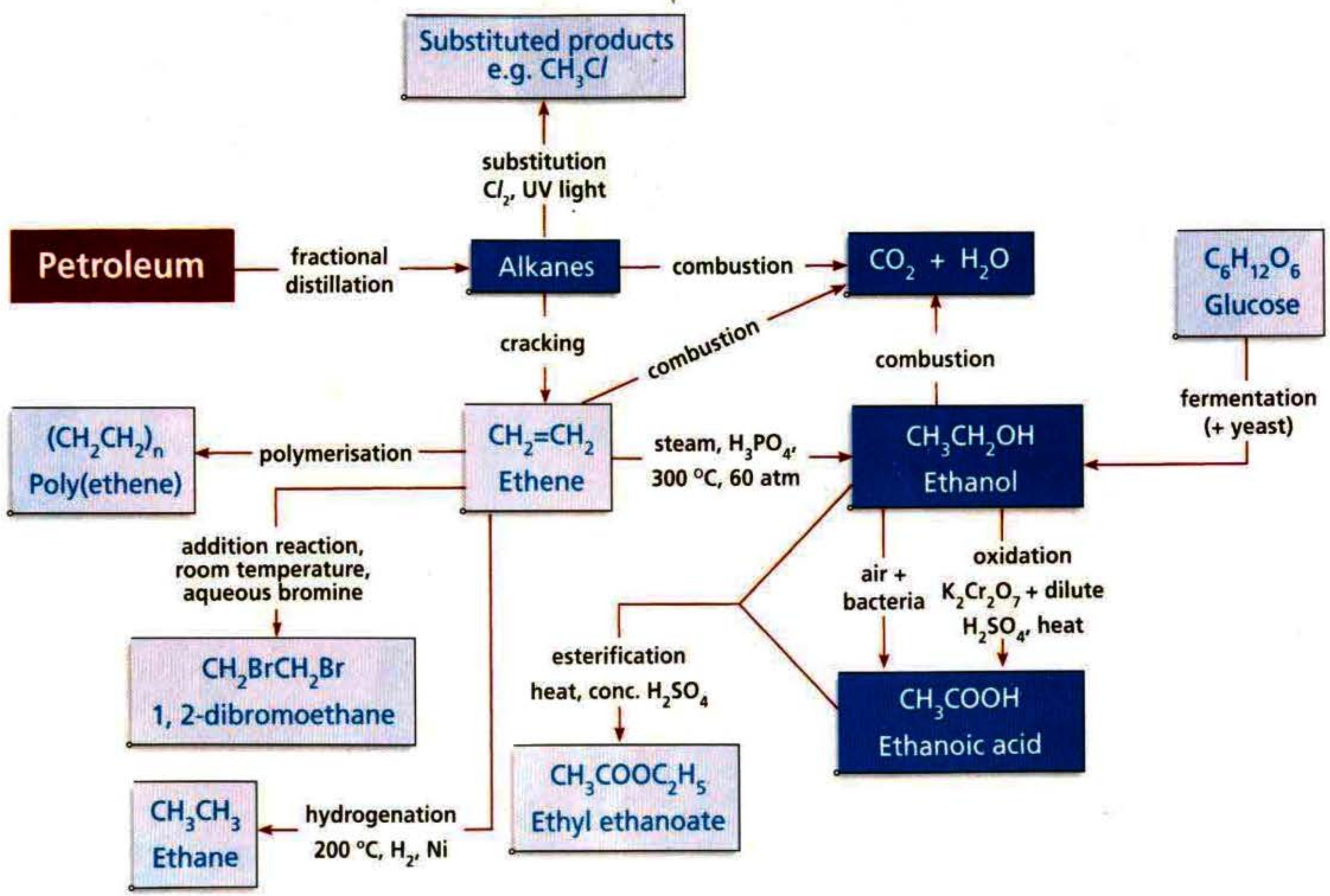
Kevlar is stronger than steel. It is used for making bullet-proof vests and army helmets. Because it is very stable at very high temperatures, it is used in the protective clothing worn by firefighters.

CRITICAL THINKING

Use the Internet or refer to organic chemistry books to find out the structural formula of Kevlar. Then answer the following:

a) What is the other monomer needed to make Kevlar?
b) What is the chemical reaction that produces Kevlar?

Inter-Conversion of Organic Compounds



Glossary

Acid. A substance that produces hydrogen ions (H^+) as the only positive ions when dissolved in water.

Activation energy. The minimum energy that molecules must possess during their collisions in order for a chemical reaction to occur.

Addition reaction. A reaction in which a molecule (element or compound) adds to an unsaturated compound to form a single new compound.

Alcohol. An organic compound containing the hydroxyl group, –OH.

Alkali. A base that is soluble in water.

Alkali metals. The elements in Group I of the Periodic Table.

Alkane. Hydrocarbon having the general formula C_nH_{2n+2}.

Alkene. Hydrocarbon that contains one or more carbon-carbon double bonds. Alkenes with only one carbon-carbon double bond have the general formula C_nH_{2n}.

Alloy. A mixture of a metal with non-metals or other metals.

Anhydrous. Anhydrous salts are salts without water of crystallisation.

Anion. A negatively charged ion which moves towards the anode during electrolysis.

Anode. A positively charged electrode in an electrolytic cell.

Aqueous. Describing the solution of a substance in water, i.e. the aqueous solution. In chemical equations, aqueous solutions are represented by the symbol (aq).

Atom. The smallest particle of an element.

Avogadro's constant. The number of particles in one mole of a substance. Its value is 6×10^{23}.

Avogadro's law. At constant temperature, the volume of a gas is directly proportional to the number of moles of the gas present.

Base. A substance that reacts with an acid to form a salt and water only.

Boiling point. The temperature at which a liquid turns rapidly to its vapour.

Carboxylic acid. An organic acid containing the carboxyl group, –COOH.

Cathode. A negatively charged electrode in an electrolytic cell.

Cation. A positively charged ion which moves towards the cathode during electrolysis.

Chromatography. A method of separating the components in a mixture.

Collision Theory. A chemical reaction can occur only if the reacting particles collide with one another.

Combustion. The chemical name for burning. Burning occurs when a substance reacts very rapidly with oxygen.

Compound. A substance formed in a chemical change when two or more elements are joined together.

Condensation. The process by which a vapour or a gas turns to a liquid on cooling.

Corrosion. The wearing away of the surface of a metal by chemical action.

Covalent bond. The type of bond formed when electrons are shared between two atoms.

Cracking. The breaking down of long chain hydrocarbon molecules with heat and/or catalyst to produce smaller hydrocarbon molecules and/or hydrogen.

Decomposition. A chemical reaction that results in the breaking down of a compound into two or more components.

Diatomic molecule. A molecule that consists of two atoms.

Displacement reaction. A reaction in which an atom or molecule takes the place of another atom or molecule in a compound.

Distillation. A process of obtaining the pure solvent from a solution. When the solution is boiled, the solvent is vaporised and the vapour condenses to reform the pure liquid.

Electrode. A rod or a plate which carries electricity in or out of an electrolyte during electrolysis.

Electrolysis. A process in which electrical energy is used to cause a chemical reaction to occur.

Electron. A negatively charged sub-atomic particle that surrounds the nucleus of an atom.

Electronic configuration. The arrangement of electrons in the various shells of an atom or a molecule.

Element. A substance made from only one type of atom. It cannot be separated into simpler substances by chemical processes or by electricity.

Endothermic reaction. A reaction which absorbs heat from the surroundings.

Evaporation. The process by which a liquid changes to its vapour on the surface of the liquid.

Exothermic reaction. A process that gives off heat to the surroundings.

Fermentation. The conversion of glucose by micro-organisms such as yeast into ethanol and carbon dioxide.

Filtrate. The clear liquid which passes through the filter during filtration.

The process of separating a solid from a liquid or a solution.

Fossil fuels. Fuels produced many millions of years ago from the decaying remains of animals or plants.

Fractional distillation. A process that separates the components in a mixture on the basis of their different boiling points. The component with the lowest boiling point boils off first and is distilled over.

Freezing point. The temperature at which a liquid changes to a solid.

Fuel. A substance that burns easily to produce energy.

Functional group. An atom or group of atoms that gives characteristic properties to an organic compound.

Giant structure. A three-dimensional network of atoms or ions packed together in a regular pattern.

Group. A vertical column of elements in the Periodic Table.

Halogen. The non-metallic elements in Group VII of the Periodic Table.

Homologous series. A family of organic compounds with members of the family having the same functional group and similar chemical properties.

Hydrated salts. Salts that contain water of crystallisation.

Hydrocarbons. Organic compounds made up from the elements hydrogen and carbon only.

Hydrogenation. The addition of a hydrogen molecule across a double bond.

Immiscible. Two liquids that do not mix.

Indicators. Compounds that have distinctly different colours in acidic and alkaline solutions.

Ion. A positively or negatively charged particle. It is formed when an atom or group of atoms loses or gains electrons.

Ionic bond. The electrostatic force that holds positive and negative ions together in an ionic compound.

Isotopes. Atoms of the same element that have the same atomic number but different nucleon numbers.

Melting point. The temperature at which a solid changes to a liquid.

Metal. An element that is shiny and conducts electricity in the solid state. Metals burn in oxygen to form basic oxides or amphoteric oxides.

Mixture. A substance made by mixing other substances together. The components in a mixture can be easily separated by physical methods.

Mole. The amount of a substance which contains 6×10^{23} particles.

Molecule. A group of atoms held together by covalent bonds. Molecules may be elements or compounds.

Nucleon number. The sum of the number of protons and neutrons in the nucleus of an atom.

Neutralisation. The reaction between an acid and a base to produce a salt and water only.

Neutron. A sub-atomic particle in the nucleus of an atom. It has a mass but no electrical charge.

Organic chemistry. The branch of chemistry that deals with carbon compounds.

A reaction where a substance gains oxygen or loses hydrogen. Oxidation is also defined as the loss of electron(s) or the increase in the oxidation state of the element.

Oxides. Compounds of an element with oxygen.

Oxidising agent.. A substance that brings about oxidation. It is itself reduced. An oxidising agent is an acceptor of electrons.

Period. A horizontal row of elements in the Periodic Table.

Periodic Table. A table that contains horizontal rows and vertical columns of elements. The elements are arranged in order of their atomic numbers and in accordance with their chemical properties.

pH scale. A scale that measures the acidity or alkalinity of a solution.

Pollution. The presence in the environment of toxic substances which are harmful to living things.

Polymer. A very large molecule built up of a number of repeating units called monomers.

Polymerisation. A chemical reaction in which simple molecules, called monomers, react with each other to form large molecules called polymers.

Polyunsaturated. Vegetable oils that contain many carbon-carbon double bonds in their molecules.

Precipitate. An insoluble solid that is produced in a solution as a result of a chemical reaction.

Protein. A polymer of amino acids.

Positively charged sub-atomic particles found in the nucleus of an atom.

Proton number. The number of protons in the nucleus of an atom.

Pure substance. A single substance which is not mixed with other substances. It has definite melting and boiling points.

Reactivity series. A list of elements in order of their reactivity. The more reactive the element, the higher its position in the series. An element at the top of the series will displace a less reactive one from a solution of its salt.

Redox reaction. A reaction where both oxidation and reduction take place simultaneously.

Reducing agent. A substance that brings about reduction. It is itself oxidised. A reducing agent is a donor of electrons.

Reduction. The removal of oxygen, the addition of hydrogen, the gain of electrons, or the decrease in the oxidation state of the substance.

Relative atomic mass. The number of times the mass of one atom of an element is heavier than $\frac{1}{12}$ of the mass of a carbon-12 atom.

Relative molecular mass. The sum of the relative atomic masses of each of the atoms in one molecule of a substance.

Residue. The solid which remains on the filter paper after filtration.

Respiration. The slow combustion of food in the cells of living organisms.

Rusting. The slow oxidation of iron in the presence of air and water to form hydrated iron(III) oxide (rust).

Salt. The ionic compound formed by the replacement of one or more hydrogen ions of an acid by a metallic ion or an ammonium ion.

Saturated hydrocarbons. Hydrocarbons that contain only single bonds between carbon atoms.

Solute. The substance that dissolves in a solvent to form a solution.

Solvent. The liquid in which a solute dissolves.

Steel. An alloy of iron and carbon.

Structural formula. A formula which shows how the atoms are arranged in a molecule.

Sublimation. The process of changing from the solid state directly to the gaseous state without passing through the liquid state.

Suspension. A mixture of a liquid and an insoluble solid where the insoluble solid remains suspended throughout the solution.

Titration. The gradual addition of a solution from a burette to another solution in a conical flask until the chemical reaction between the two solutions is complete.

Unsaturated molecule. Any hydrocarbon that contains one or more carbon-carbon double bonds.

Valence electrons. Electrons in the outer shell that are used by the atom for forming chemical bonds.

Water of crystallisation. Water molecules that are chemically bonded in the crystals of some salts.

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The Periodic Table of the Elements

I	II											III	IV	V	VI	VII	0
							1 H Hydrogen 1										4 He Helium 2
7 Li Lithium 3	9 Be Beryllium 4											11 B Boron 5	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	20 Ne Neon 10
23 Na Sodium 11	24 Mg Magnesium 12											27 A*l* Aluminium 13	28 Si Silicon 14	31 P Phosphorus 15	32 S Sulphur 16	35.5 C*l* Chlorine 17	40 Ar Argon 18
39 K Potassium 19	40 Ca Calcium 20	45 Sc Scandium 21	48 Ti Titanium 22	51 V Vanadium 23	52 Cr Chromium 24	55 Mn Manganese 25	56 Fe Iron 26	59 Co Cobalt 27	59 Ni Nickel 28	64 Cu Copper 29	65 Zn Zinc 30	70 Ga Gallium 31	73 Ge Germanium 32	75 As Arsenic 33	79 Se Selenium 34	80 Br Bromine 35	84 Kr Krypton 36
85 Rb Rubidium 37	88 Sr Strontium 38	89 Y Yttrium 39	91 Zr Zirconium 40	93 Nb Niobium 41	96 Mo Molybdenum 42	99 Tc Technetium 43	101 Ru Ruthenium 44	103 Rh Rhodium 45	106 Pd Palladium 46	108 Ag Silver 47	112 Cd Cadmium 48	115 In Indium 49	119 Sn Tin 50	122 Sb Antimony 51	128 Te Tellurium 52	127 I Iodine 53	131 Xe Xenon 54
133 Cs Caesium 55	137 Ba Barium 56	139 La Lanthanum 57	178 Hf Hafnium 72	181 Ta Tantalum 73	184 W Tungsten 74	185 Re Rhenium 75	190 Os Osmium 76	192 Ir Iridium 77	195 Pt Platinum 78	197 Au Gold 79	201 Hg Mercury 80	204 T*l* Thallium 81	207 Pb Lead 82	209 Bi Bismuth 83	Po Polonium 84	At Astatine 85	Rn Radon 86
Fr Francium 87	226 Ra Radium 88	227 Ac Actinium 89															

58-71 Lanthanoid series
90-103 Actinoid series

140 Ce Cerium 58	141 Pr Praseodymium 59	144 Nd Neodymium 60	Pm Promethium 61	150 Sm Samarium 62	152 Eu Europium 63	157 Gd Gadolinium 64	159 Tb Terbium 65	162 Dy Dysprosium 66	165 Ho Holmium 67	167 Er Erbium 68	169 Tm Thulium 69	173 Yb Ytterbium 70	175 Lu Lutetium 71
232 Th Thorium 90	Pa Protactinium 91	238 U Uranium 92	Np Neptunium 93	Pu Plutonium 94	Am Americium 95	Cm Curium 96	Bk Berkelium 97	Cf Californium 98	Es Einsteinium 99	Fm Fermium 100	Md Mendelevium 101	No Nobelium 102	Lr Lawrencium 103

Key

a X b	a = relative atomic mass X = atomic symbol b = proton (atomic) number

Name ________________________________

STOICHIOMETRY WORKSHEET (MOLE-MOLE)

1. Magnesium reacts with hydrochloric acid according to the following balanced chemical equation:

$$Mg\ (s) + 2\ HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2\ (g)$$

   If two moles of hydrochloric acid react with excess magnesium, how many moles of hydrogen gas will be produced?

2. Aluminum reacts with HCl to produce aluminum chloride and hydrogen gas. Write a balanced equation for the reaction and calculate the number of moles of HCl required to react with 0.87 moles of Al.

3. Glucose ($C_6H_{12}O_6$) combines with O_2 in the body to produce carbon dioxide and water. Write a balanced equation for this reaction. How many moles of O_2 are required to combine with 0.25 moles of glucose? How many moles of CO_2 and H_2O would be produced in this reaction?

4. Calcium carbonate combines with HCl to produce calcium chloride, water, and carbon dioxide gas. Write the balanced equation for this reaction. How many moles of HCl are required to react with 2.5 moles of calcium carbonate? How many moles of carbon dioxide would be produced?

5. Zinc reacts with sulfuric acid (H_2SO_4) to yield zinc sulfate and hydrogen gas. How many moles of hydrogen will be produced if 0.36 moles of zinc react with an equal amount of H_2SO_4?

Stoichiometry Worksheet

1. Na_2SIO_3 (s) + 8 HF(aq) → H_2SiF_6 (aq) + 2 NaF(aq) + 3 H_2O (l)

a. How many moles of HF are needed to react with 0.300 mol of Na_2SiO_3?

b. How many grams of NaF form when 0.500 mol of HF reacts with excess Na_2SiO_3?

c. How many grams of Na_2SiO_3 can react with 0.800 g of HF?

2. $C_6H_{12}O_6$ (aq) → 2 C_2H_5OH (aq) + 2 CO_2 (g)

a. How many moles of CO_2 are produced when 0.400 mol of $C_6H_{12}O_6$ reacts in this fashion?

b. How many grams of $C_6H_{12}O_6$ are needed to form 7.50 g of C_2H_5OH?

c. How many grams of CO_2 form when 7.50 g of C_2H_5OH are produced?

3. Fe_2O_3 (s) + CO (g) → Fe(s) + CO_2 (g) (unbalanced!)

a. Calculate the number of grams of CO that can react with 0.150 kg of Fe_2O_3

b. Calculate the number of grams of Fe and the number of grams of CO_2 formed when 0.150 kg of Fe_2O_3 reacts

4. 2 NaOH (s) + CO_2 (g) $\rightarrow$ Na_2CO_3 (s) + H_2O (l)

a. Which reagent is the limiting reactant when 1.85 mol NaOH and 1.00 mol CO_2 are allowed to react?

b. How many moles of Na_2CO_3 can be produced?

5. C_6H_6 + Br_2 $\rightarrow$ C_6H_5Br + HBr

a. What is the theoretical yield of C_6H_5Br in this reaction when 30.0 g of C_6H_6 reacts with 65.0 g or Br_2?

b. If the actual yield of C_6H_5Br was 56.7 g, what is the percent yield?